CHEMISTRY (EM)

9TH CLASS

Presented by:

Urdu Books Whatsapp Group

STUDY GROUP

0333-8033313
راؤ ایاز

0343-7008883
پاکستان زندہ باد

0306-7163117
محمد سلمان سلیم

Unit 1: Fundamentals of Chemistry
Exercise Questions

Exercise Multiple Choice Question Answers

1. **Industrial chemistry deals with the manufacturing of compounds:**
(a) in the laboratory (b) on micro scale
(c) on commercial scale (d) on economic scale
2. **Which one of the following can be separated by physical means?**
(a) mixture (b) element (c) compound (d) radical
3. **The most abundant element occurring in the oceans is:**
(a) oxygen (b) hydrogen (c) nitrogen (d) silicon
4. **Which one of the following element is found in most abundance in the earth's crust?**
(a) oxygen (b) aluminum (c) silicon (d) iron
5. **The third abundant gas found in the atmosphere is:**
(a) carbon monoxide (b) oxygen (c) nitrogen (d) argon
6. **One amu (atomic mass unit) is equivalent to:**
(a) 1.66×10^{-24} mg (b) 1.66×10^{-24}g (c). 1.66×10^{-21}g (d) 1.66×10^{-23}g
7. **All of the followings are tri-atomic molecule except:**
(a) H_2 (b) O_3 (c) H_2O (d) CO_2
8. **The mass of one molecule of water is:**
(a) 18 amu (b) 18 g (c) 18 mg (d) 18 kg
9. **The molar mass of H_2S0_4 is·**
(a) 98 g (b) 48 amu (c) 4.8 g (d) 98 amu
10. **Molar mass is usually expressed in grams. Which one of the following is molar mass of O_2 in amu?**
(a) 32 amu (b) 53.2×10^{-24} amu
(c) 1.92×10^{-25} amu (d) 192.64×10^{-25} amu
11. **How many numbers of moles are equivalent to 8 grams of CO_2?**
(a) 0.15 (b) 0.18 (c) 0.21 (d) 0.24
12. **Which one of the following pairs has the same number of ions?**
(a) 1 mole of NaCl and 1 mole of $MgCl_2$
(b) 1/2 mole of NaCl and 1/2 mole of $MgCl_2$
(c) 1/2 mole of NaCl and 1/3 mole of $MgCl_2$
(d) 1/3 mole of NaCl and 1/2 mole of $MgCl_2$
13. **Which one of the following pairs has the same mass?**
(a) 1 mole of CO and 1 mole of N_2 (b) 1 mole of CO and 1 mole of CO_2
(c) 1 mole of O_2 and 1 mole of N_2 (d) 1 mole of O_2 and 1 mole of CO_2

ANSWR KEY

1	c	3	a	5	d	7	a	9	a	11	b	13	a
2	a	4	a	6	b	8	a	10	a	12	c		

Exercise Short Questions Answers

Q.1 Define industrial chemistry and analytical chemistry.

Ans: **Industrial Chemistry**

"The branch of chemistry that deals with the manufacturing of chemical substances (elements and compounds) on commercial scale, is called industrial chemistry."

Applications:

i. It deals with the manufacturing of basic chemicals such as oxygen, chlorine, ammonia, caustic soda, nitric acid and sulphuric acid.

ii. Use of these chemicals to provide the raw materials for many other industries such as fertilizers, soap, textiles, agricultural products, paints and paper etc

Analytical Chemistry

"The branch of chemistry that deals with separation and analysis of a sample to identify its components is called analytical chemistry. The separation is carried out prior to qualitative and quantitative analysis."

Qualitative Analysis

It provides the identity of a substance (composition of chemical species).

Quantitative Analysis

It determines the amount of each component present in the sample.

Scope

In this branch different techniques and instruments used for analysis are studied.

The scope of this branch covers food, water, environmental and clinical analyses.

Q.2 How can you differentiate between organic and inorganic chemistry?

Ans: **Organic Chemistry**

"The branch of chemistry that deals with the study of covalent compounds of carbon and hydrogen and their derivatives is called organic chemistry."

Scope

Organic chemists determine the structure and properties of these naturally occurring as well as synthesized compounds.

Scope of this branch covers petroleum, petrochemicals and pharmaceutical industries.

ختم نبوت صلی اللہ علیہ وسلم زندہ باد　　　　　عظمت صحابہ زندہ باد

السلام علیکم ورحمۃ اللہ وبرکاتہ:

معزز ممبران: آپ کا وٹس ایپ گروپ ایڈمن "**اردو بکس**" آپ سے مخاطب ہے۔

آپ تمام ممبران سے گزارش ہے کہ:

- گروپ میں صرف PDF کتب پوسٹ کی جاتی ہیں لہذا کتب کے متعلق اپنے کمنٹس / ریویوز ضرور دیں۔ گروپ میں بغیر ایڈمن کی اجازت کے کسی بھی قسم کی (اسلامی و غیر اسلامی، اخلاقی، تحریری) پوسٹ کرنا سختی سے منع ہے۔
- گروپ میں معزز، پڑھے لکھے، سلجھے ہوئے ممبرز موجود ہیں اخلاقیات کی پابندی کریں اور گروپ رولز کو فالو کریں بصورت دیگر معزز ممبرز کی بہتری کی خاطر ریموو کر دیا جائے گا۔
- کوئی بھی ممبر کسی بھی ممبر کو انباکس میں میسیج، مس کال، کال نہیں کرے گا۔ رپورٹ پر فوری ریموو کر کے کاروائی عمل میں لائے جائے گی۔
- ہمارے کسی بھی گروپ میں سیاسی و فرقہ واریت کی بحث کی قطعاً کوئی گنجائش نہیں ہے۔
- اگر کسی کو بھی گروپ کے متعلق کسی قسم کی شکایت یا تجویز کی صورت میں ایڈمن سے رابطہ کیجئے۔
- سب سے اہم بات:

گروپ میں کسی بھی قادیانی، مرزائی، احمدی، گستاخِ رسول، گستاخِ امہات المؤمنین، گستاخِ صحابہ و خلفائے راشدین حضرت ابو بکر صدیق، حضرت عمر فاروق، حضرت عثمان غنی، حضرت علی المرتضیٰ، حضرت حسنین کریمین رضوان اللہ تعالیٰ اجمعین، گستاخ اہلبیت یا ایسے غیر مسلم جو اسلام اور پاکستان کے خلاف پراپیگنڈا میں مصروف ہیں یا ان کے روحانی و ذہنی سپورٹرز کے لئے کوئی گنجائش نہیں ہے لہذا ایسے اشخاص بالکل بھی گروپ جوائن کرنے کی زحمت نہ کریں۔ معلوم ہونے پر فوراً ریموو کر دیا جائے گا۔

- تمام کتب انٹرنیٹ سے تلاش / ڈاؤنلوڈ کر کے **فری آف کاسٹ** وٹس ایپ گروپ میں شیئر کی جاتی ہیں۔ جو کتاب نہیں ملتی اس کے لئے معذرت کر لی جاتی ہے۔ جس میں محنت بھی صَرف ہوتی ہے لیکن ہمیں آپ سے صرف دعاؤں کی درخواست ہے۔
- عمران سیریز کے شوقین کیلئے علیحدہ سے عمران سیریز گروپ موجود ہے۔
- **لیڈیز کے لئے الگ گروپ کی سہولت موجود ہے جس کے لئے ویریفکیشن ضروری ہے۔**
- اردو کتب / عمران سیریز یا سٹڈی گروپ میں ایڈ ہونے کے لئے ایڈمن سے وٹس ایپ پر بذریعہ میسیج رابطہ کریں اور جواب کا انتظار فرمائیں۔ برائے مہربانی اخلاقیات کا خیال رکھتے ہوئے موبائل پر کال یا ایم ایس کرنے کی کوشش ہر گز نہ کریں۔ ورنہ گروپس سے تو ریموو کیا ہی جائے گا بلاک بھی کیا جائے گا۔

نوٹ: ہمارے کسی گروپ کی کوئی فیس نہیں ہے۔ سب فی سبیل اللہ ہے

0306-7163117	0343-7008883	0333-8033313
محمد سلمان سلیم	پاکستان زندہ باد	راؤ ایاز

پاکستان زندہ باد　　　　　پاکستان پائندہ باد

اللہ تبارک تعالیٰ ہم سب کا حامی و ناصر ہو

Inorganic Chemistry

"The branch of chemistry that deals with the study of all elements and their compounds except those of compounds of carbon and hydrogen (hydrocarbons) and their derivatives is called inorganic chemistry."

Applications

It has applications in every aspect of the chemical industry such as glass, cement, ceramics and metallurgy (extraction of metals from ores).

Q.3 Give the scope of biochemistry.

Ans: *"The branch of chemistry that deals with the study of structure, composition, and chemical reactions of substances found in living organism."*

Scope

It covers all chemical processes taking place in living organisms such as synthesis and metabolism of bio-molecules like carbohydrates, proteins, fats etc.

Emergence of biochemistry as a separate discipline

Biochemistry emerged as a separate discipline when scientists began to study:

i. How living things obtain energy from food?

ii. How' the fundamental biological changes occur during a disease?

Applications:

It is applied in the fields of medicine, food science and agriculture.

Q.4 Wow does homogeneous mixture differ from heterogeneous mixture?

Ans:

Homogeneous Mixture	Heterogeneous Mixture
Mixtures that have uniform composition throughout are called homogeneous mixtures. *It is called solution.*	*Those mixtures in which composition is not uniform throughout are called heterogeneous mixtures.*
For example: Air, gasoline and ice cream	**For example:** Soil, rock, wood, concrete and paint etc.

Q.5 What is the relative atomic mass? How it is related to gram?

Ans: Relative atomic mass

"The average mass of an atom of an element as compared to 1/12th (one-twelfth) the mass of one atom of carbon-12 isotope is called relative atomic mass."

Unit of relative atomic mass:

Its unit is atomic mass unit, with symbol amu.

Atomic mass unit:

"One atomic mass unit is 1/12th the mass of one atom of carbon-12th."

When this atomic mass unit is expressed in grams it is.

$$1 \text{ amu} = 1.66 \times 10^{-24} \text{ g}$$

Q.6 Define empirical formula with example.

Ans: **Empirical Formula:**

"It is the simplest whole number ratio of atoms present in a compound."

The empirical formula of a compound is determined by knowing the percentage composition of a compound.

Example:

Glucose has simplest ratio 1: 2: 1 of carbon, hydrogen and oxygen respectively. Hence its empirical formula is CH_2O.

Q.7 State three reason why do you think air is a mixture and water is a compound?

Ans: **Reasons:**

i. Water is a compound because it is formed by chemical combination of hydrogen and oxygen whereas air is formed by simple mixing of different gases.

ii. Water has fixed ratio between masses of hydrogen and oxygen, whereas in air ratio between masses of component gases is not fixed.

iii. Water has definite melting and boiling points whereas air does not have any fixed melting and boiling point.

Q.8 Explain why are hydrogen and oxygen considered elements whereas water as a compound.

Ans: Hydrogen and oxygen are elements because they have same type of atoms, having same atomic number and it cannot be decompose into simple substances by chemical means. Water is considered as compound because it is a substance made up of two or more elements chemically combined together in a fixed ratio by mass. As a result of this combination oxygen and hydrogen lose their own properties and produce new substance (H_2O).

Q.9 What is the significance of the symbol of an element?

Ans: **Significance of the symbol of an element:**

Symbols are used for elements instead of writing of their complete names. So, it takes less time/save time and element can be recognized by that symbol in all over the world.

i. Symbol represents the name of an element.
ii. It represents one atom of the element
iii. It represents one mole of atoms of the element.
iv. It represents atomic mass of an element.
v. It helps to write and understand chemical equation for different chemical reactions.
vi. Periodic table is based on symbols of different elements.

For example: Oxygen (O), Sulphur (S), Nitrogen (N)

Q.10 State the reasons: soft drink is a mixture and water is a compound.

Ans:

Mixture (Soft Drink)	Compound (Water)
• Soft drink is made up of simple mixing up of substances without any fixed ratio.	• Water is formed by chemical combination of atoms of elements hydrogen and oxygen in a fixed ratio of 1:8 by mass.
• Soft drink has heterogeneous composition.	• Water has homogeneous composition.
• Its components can be separated by physical means.	• Its components can't be separated by physical means

Q.11 Classify the following into element, compound and mixture:

- **He and H_2**
- **CO and CO_2**
- **Water and milk**
- **Gold and brass**
- **Iron and steel**

Ans:

(i) He and H_2: He and H_2 are elements

(ii) CO and Co: CO is a compound and Co is an element.

(iii) Water and milk: Water is a compound and milk is a mixture

(iv) Gold and brass: Gold is an element and brass is a mixture

(v) Iron and steel: Iron is an element and steel is a mixture.

Q.12 Define atomic mass unit. Why is it needed?

Ans: **Atomic mass unit**

"The mass equal to one twelfth of the mass of a carbon -12 atom is called atomic mass unit."

The atomic mass unit is abbreviated as amu.

1 amu = 1/12 × mass of C-12 atom

The mass of one atom of carbon -12 is 12 amu.

It is the unit used for the relative atomic mass. It is used to compare masses of atoms.

Q.13 State the nature and name of the substance formed by combining the following:
i. Zinc + Copper ii. Water + Sugar iii. Aluminium + Sulphur
iv. Iron + Chromium + Nickel

Ans:

(i) Zinc + Copper

It is a mixture or alloy. The name of alloy is brass.

(ii) Water + Sugar

It is a mixture. The name of mixture or solution is syrup.

(iii) Aluminium + Sulphur

It forms compound. The name of compound is aluminium -sulphide.

(iv) Iron + Chromium + Nickel

It is a mixture or alloy. The name of alloy is nichrome.

Q.14 Differentiate between molecular mass and formula mass, which of the following will be molecular formula?

- **H_2O**
- **NaCl**
- **KI**
- **H_2SO_4**

Ans:

Molecular Mass	Formula Mass
i. The sum of atomic masses of all the atoms present in one molecule of a molecular substance called molecular mass. **ii.** The term molecular mass is used for compounds that exist as molecules.	**i.** *The sum of atomic masses of all the atoms present in one formula unit of an ionic compound is called formula mass.* **ii.** The term formula mass is used for compounds that exits as formula units i.e. the compounds consists of ions.
For example: Molecular mass of water is 18 amu and that of carbon is 44 amu	**For example:** Formula mass of sodium chloride is 58.5 amu and that of $CaCO_3$ is 100 amu.

H_2O and H_2SO_4 are the molecular formula.

Q.15 Which has more then atoms: 10 g of Al or 10 g of Fe?

Ans: 10 g of Al has more atoms than 10 g of Fe.

For Al

i. Given mass of Al = 10g

Molar mass of Al = 27 g mol^{-1}

No of atoms in 10g of Al = No of moles × N_A

$$\text{Number of atom} = \frac{\text{Mass}}{\text{Molar Mass}} \times N_A$$

$$= \frac{10}{23} \times 6.02 \times 10^{23}$$

$$= 2.23 \times 10^{23} \text{ atoms}$$

ii. Given mass of Fe = 10g

Molar mass of Fe = 56 g mol^{-1}

For Fe

$$\text{Number of atom} = \frac{\text{Mass}}{\text{Molar Mass}} \times N_A$$

$$= \frac{10}{56} \times 6.02 \times 10^{23}$$

$$= 1.075 \times 10^{23}$$

Therefore: Aluminium has more number of atoms than iron.

Q.16 Which one has more molecules: 9 g of water or 9 g of sugar ($C_{12}H_{22}O_{11}$)?

Ans: 9 g of H_2O has more molecules than 9 g of $C_6H_{12}O_6$

i. Given mass of water (H_2O) = 9 g

Molar mass of water (H_2O) = 18 g mol^{-1}

Number of molecules in 9g of water $= \frac{\text{Given Mass}}{\text{Molar Mass}} \times N_A$

$$= \frac{9}{18} \times 6.02 \times 10^{23} \text{ molecules}$$

$$= 3.01 \times 10^{23} \text{ molecules}$$

ii. Given mass of sugar = 9 g

Molar mass of sugar = 342 g mol^{-1}

Number of molecules in 9g of sugar $= \frac{\text{Mass}}{\text{Molar Mass}} \times N_A$

$$= \frac{9}{342} \times 6.02 \times 10^{23}$$

$$= 1.584 \times 10^{22} \text{ molecules}$$

Therefore: 9 g of H_2O has more molecules than 9 g of $C_{12}H_{22}O_1$.

Q.17 Which one has more formula units: 1 g of Na*Cl* or 1 g of K*Cl*?

Ans:

i. Given mass = 1g

Formula mass of NaCl = 58.5 g mol^{-1}

Formula units in 1g of NaCl $= \frac{\text{Given mass}}{\text{Formula mass}} \times N_A$

$$= \frac{1}{58.5} \times 6.02 \times 10^{23}$$

$$= 1-029 \times 10^{22} \text{ formula units}$$

ii. Given mass of KCl = 1g

Formula mass of KCl = 74.5 g mol^{-1}

Formula units $= \frac{\text{Given mass}}{\text{Formula mass}} \times N_A$

$$= \frac{1}{74.5} \times 6.02 \times 10^{23}$$

$$= 8.080 \times 10^{21} \text{ formula units}$$

Therefore 1g of NaC*l* has more formula units than 1g of KC*l*.

Q.18 Differentiate between homoatomic and heteroatomic molecules with examples.

Ans:

Homoatomic molecules	Heteroatomic molecules
A molecule containing same type of atoms is called homoatomic molecule.	*A molecule consists of different kinds of atoms is called a heteroatomic molecule.*
Examples: Hydrogen (H_2), Oxygen (O_3) and sulphur (S_8) These are molecules of elements	**Examples:** CO_2, H_2O and NH_3 These are molecules of compounds

Q.19: In which one of the followings the number of hydrogen atoms in more? 2 moles of HCl or 1 mole of NH_3 (Hint: 1 mole of a substance contains as much number of moles of atoms as are in 1 molecule of a substance.)

Ans: No of moles of hydrogen in 1 mole of HCl = 1 mole

No of moles of hydrogen in 2 moles of HCl = 2 moles

Whereas No of moles of hydrogen in 1 mole of NH_3= 3 moles

Hence 1 mole of NH_3 contains 3 moles of hydrogen will have more hydrogen atoms than 2 moles of hydrogen present in 2 moles of HCl.

Exercise Long Question Answers

Q.1 Define element and classify the elements with examples.

Ans: See Q. No. **4** (Subjective Part, Long Questions Answers)

Q.2 List five characteristics by which compounds can be distinguished from mixtures.

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.3 Differentiate between the following with examples:

(a) Molecule and gram molecule **(b) Atom and gram atom**

(c) Molecular mass and molar mass **(d) Chemical formula and gram formula**

Ans: See Q. No. **21** (Subjective Part, Long Questions Answers)

Q.4 Mole is SI unit for the amount of a substance. Define it with examples?

Ans: See Q. No. **23** (Subjective Part, Long Questions Answers)

Exercise Solved Numericals

Q.1 Sulphuric acid is the king of chemicals. If you need 5 moles of sulphuric acid for a reaction, how many grams of it will you weigh?

Given Data:

Number of moles of H_2SO_4 = 5 moles

Molar mass of H_2SO_4 = $2(1)+1(32)+4(16)$

= 2 + 32 + 64

= 98 g/mol

To find:

Mass of H_2SO_4 = ?

Solution:

Number of Moles = $\frac{\text{Mass of } H_2SO_4}{\text{Molar mass of } H_2SO_4}$

Mass in grams = No of moles × molar mass = 5×98 = 490 g

Therefore, 5 moles of sulphuric acid will have mass 490 g.

Q.2 Calcium carbonate is insoluble in water. If you have 40 g of it; how m~ Ca^{2+} and CO_3^{2-} ions are present in it?

Given Data:

Mass of calcium carbonate = 40g

Molar mass of calcium carbonate = $CaCO_3$

= (40×1) + (12×1)+(16×3)

= 40 + 12 + 48

= 100 g/mol

To find:

Number of Ca^{2+} ions = ?

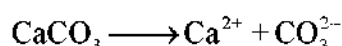
Number of CO_3^{2-} ions = ?

Solution:

Number of Moles of $CaCO_3$ = $\frac{\text{Mass}}{\text{Molar mass}}$

$= \frac{40}{100} = 0.4$ mol

Balanced equation:

$$CaCO_3 \longrightarrow Ca^{2+} + CO_3^{2-}$$

Number of moles of $CaCO_3$ = 0.4 mole

Number of moles of Ca^{2+} ions in one mole of $CaCO_3$ = 6.02×10^{23}

Number of Ca+2 ion in 0.4 moles of $CaCO_3$ = No of moles × N_A

= $0.4\times6.02\times10^{23}$

= 2.408×10^{23} ions

We know that

No of Ca^{2+} ions = No of CO_3^{2-} ions

No of CO_3^{2-} = 2.40×10^{23} ions

= 2.408×10^{23} ions

Q.3 If you have 6.02×10^{23} ions of aluminium; how many sulphate ions will be required to prepare $Al_2(SO_4)_3$?

Given Data:

Number of ions of Al^{3+} = 6.02×10^{23}

Number of sulphate ions in $Al_2(SO_4)_3$ =?

Solution:

$$Al^{3+} + 3SO_4^{2-} \longrightarrow Al_2(SO_4)_3$$

According to balanced chemical equation:

No of moles of SO_4^{2-} ion required for 2 moles of Al^{3+} ions = 3

No of moles of SO_4^{2-} ions for 1 mole of Al^{3+} = 3/2 = 1.5 moles

Thus, number of SO_4^{2-} ions = No of moles × N_A

= $1.5\times6.02\times10^{23}$

= 9.03×10^{23} ions

Q.4 Calculate the number of molecules of the following compounds:

a. 16 g of H_2CO_3 b. 20 g of HNO_3 c. 30 g of $C_6H_{12}O_6$

Given Data:

a. 16g of H_2CO_3

Given mass of H_2CO_3 = 16g

Molar mass of H_2CO_3 = 2(1)+1(12)+3(16)

= 2 + 12 + 48

= 62g / mol

Number of molecules of H_2CO_3 = ?

Number of molecules of H_2CO_3 = $\frac{\text{Givenmass of } H_2CO_3}{\text{Molar mass of } H_2CO_3} \times N_A$

$= \frac{16}{62} \times 6.02\times10^{23}$

= 1.505×10^{23} molecules

b. 20g of HNO_3

Given data:

Given mass of HNO_3 = 20g

Molar mass of HNO_3 = 1(1)+1(14)+3(16)

= 1+14+48=63g/mol

Number of molecules of HNO3 = ?

Solution:

Number of molecules of HNO_3 $= \frac{\text{Given mass of } HNO_3}{\text{Molar mass of } HNO_3} \times N_A$

$= \frac{20}{63} \times 6.02 \times 10^{23}$

$= 1.91 \times 10^{23}$ molecules

= 1.908×10^{23} molecules

c. 30g of $C_6H_{12}O_6$

Given data:

Given mass of $C_6H_{12}O_6$ = 30g.

Molar mass of $C_6H_{12}O_6$ = 72 + 12 + 96

= 180g / mol

Number of moles of $C_6H_{12}O_6 = \frac{\text{Given mass of } C_6H_{12}O_6}{\text{Molar mass of } C_6H_{12}O_6} \times N_A$

$= \frac{30}{180} \times 6.02 \times 10^{23}$

= 1×10^{23} molecules

Q.5 Calculate the number of ions in the following compounds:

a. 10g of $AlCl_3$ b. 30 g of $BaCl_2$ c. 58 g of H_2SO_4

Ans:

a. 10g of $AlCl_3$

Given Data:

Given mass of $AlCl_3$ = 10g

Molar mass of $AlCl_3$ = 1(27)+3(35.5)

= 133.5 g / mol

No. of ions of $AlCl_3$ = ?

No of formula units of $AlCl_3$ in 10g = $\frac{\text{Given mass of } AlCl_3}{\text{Molar mass of } AlCl_3}$

$= \frac{10}{133.5} \times 6.02 \times 10^{23}$

$= 0.451 \times 10^{23}$ formula units

1 formula unit of $AlCl_3$ contains total number of ions = 4ions

4.51×10^{22} formula units of $AlCl_3$ contain total number of ions = $4 \times 0.451 \times 10^{23}$

= 1.80×10^{24} ions

Therefore number of ions in 10g of $AlCl_3$ = 1.80×10^{23} ions

b. 30g of $BaCl_2$

Given Data:

Given mass of $BaCl_2$ = 30g

Molar mass of $BaCl_2$ = 1(137)+2(35.5)

= 137+71=208 g/mol

No of ions of 30g of $BaCl_2$ = ?

No. of formula units in 30g of $BaCl_2$ = $\frac{\text{Givenmass of } BaCl_2}{\text{Molar mass of } BaCl_2} \times N_A$

$= \frac{30}{208} \times 6.02\times10^{23} = 0.86\times10^{23}$ formula units

1 formula unit of $BaCl_2$ contains total number of ions = 3

0.86×10^{23} formula units will contain total number of ions = $3\times0.86\times10^{23}$ ions

= 2.58×10^{23} ions

c. 58g of H_2SO_4

Given Data:

Given mass of H_2SO_4 = 58g

Molar mass of H_2SO_4 = 2(1)+1(32)+4(16)

=2+32+64

=98 g/mol

Number of ion of H_2SO_4 = ?

Number of formula units in 58g of H_2SO_4 = $\frac{\text{Given mass of } H_2SO_4}{\text{Molar mass of } H_2SO_4} \times N_A$

$= \frac{58}{98} \times 6.02\times10^{23}$

= 3.56×10^{23} formula units

1 formula unit of H_2SO_4 contains total number of ions = 3 ions

3.56×10^{23} formula units of H2SO4 contain total number of ions = $3\times3.56\times10^{23}$

= 10.682×10^{23} ions

Q.6 What will be the mass of 2.05×10^{16} molecules of H_2SO_4

Ans:

Given Data:

Number of molecules of H_2SO_4 = 2.05×10^{16}

Molar mass of H_2SO_4 = 2(1)+1(32)+4(18)=2+32+64=98g/mol

Mass of H_2SO_4 = ?

Solution:

Number of molecules of H_2SO_4 = $\frac{\text{Mass of } H_2SO_4}{\text{Molar mass of } H_2SO_4} \times N_A$

$2.05\times10^{16} = \frac{\text{Mass of } H_2SO_4}{98} \times 6.02\times10^{23}$

Mass of H_2SO_4 = $\frac{2.05\times10^{16} \times 98}{6.02\times10^{23}}$

= 3.332×10^{-6}g

Q.7 How many total atoms are required to prepare 60 g of HNO_3?

Given Data:

Given mass of HNO_3 = 60g

Molar mass of HNO_3 = 1(1)+1(14)+3(16)

= 1 + 14 + 48

= 63g / mol

No of atoms of HNO_3 = ?

Solution:

Number of molecules of HNO_3 $= \frac{\text{Given mass of } NHO_3}{\text{Molar mass of } HNO_3} \times N_A$

$= \frac{60}{63} \times 6.02 \times 10^{23}$

$= 0.95 \times 6.02 \times 10^{23}$ moles

$= 5.73 \times 10^{23}$ molecules

As one molecule of HNO_3 contain atoms = 5 atoms

There fore,

5.73×10^{23} molecules contain No of atoms $= 5 \times 5.73 \times 10^{23}$

$= 28.5 \times 10^{23}$

$= 2.87 \times 10^{24}$ atoms

Q.8 How many ions of Na^+ and Cl^{-1} will be present in 30 g of NaCl?

Given Data:

Given mass of NaCl = 30g

Molar mass of NaCl = 1(23)+3(35.5)

= 23 + 35.5

= 58.5g / mol

To Find:

No of Na+ ions = ?

No of Cl- ions = ?

Solution

Number of formula units of NaCl $= \frac{\text{Given mass of NaCl}}{\text{Molar mass of NaCl}} \times N\frac{A}{A}$

$= \frac{30}{58.5} \times 6.02 \times 10^{23}$

$= 3.08 \times 10^{23}$ formula units

As,

1 formula unit of NaCl contains number of Na^{+1} ions = 1

3.08×10^{23} formula units of NaCl contain number of Na^+ ions $= 3.08 \times 10^{23}$ ions

We knew that in NaCl

Number of Na^+ ions = number of Cl^- ion

Thus number of Cl ions $= 3.08 \times 10^{23}$

Total number of sodium ions (Na^+) and chloride ions (Cl) $= 6.16 \times 10^{23}$ ions

Q.9 How many molecules of HCI will be required to have 10 grams of it?

Given mass of HCl = 10g
Molar mass of HCl = 1(1)+1(35.5)=1+35.5 = 36.5g / mol
Number of molecules of HCl = ?

Solution:

$$\text{Number of molecules of HCl} = \frac{\text{Given mass of HCl}}{\text{Molar mass of HCl}} \times N_A$$

$$= \frac{10}{36.5} \times 6.02 \times 10^{23}$$

$$= \mathbf{1.64 \times 10^{23}} \textbf{ molecules}$$

Q.10 How many grams of Mg will have the same number of atoms as 6 grams of C have?

Given data

Given mass of carbon = 6g
Atomic mass of carbon = 12 g/ mole

To find:

Mass of Mg = ?

Solution:

$$\text{No. of moles of carbon} = \frac{\text{Given mass of Carbon}}{\text{Molar mass of Carbon}}$$

$$= \frac{6}{12}$$

$$= 0.5 \text{ mol}$$

Number of carbon atoms = Number of moles × N_A
Number of carbon atoms in 0.5 moles of carbon = $0.5 \times 6.02 \times 10^{23}$ atoms
= 3.01×10^{23} atoms

As

The number of atoms of mg and carbon are same, so their number of moles are also equal.

Thus,

$$\text{No. of atoms of Mg} = \frac{\text{Given mass of Mg}}{\text{Molar mass of Mg}} \times N_A$$

$$3.01 \times 10^{23} = \frac{\text{mass of Mg}}{24} \times 6.02 \times 10^{23}$$

$$\text{Mass of Mg} = \frac{3.01 \times 10^{23} \times 24}{6.02 \times 10^{23}}$$

$$= 12g$$

So
12g of Mg will have same number of atoms as 6g of carbon.

Last Updated: November 2020
Report any mistake at freeilm786@gmail.com

Unit 1: Fundamentals of Chemistry
Long Questions

Q.1 (a) What is meant by science?
(b) Define chemistry. Give its importance in daily life.

Ans:

(a) Science (Latin Scientia meaning knowledge)

The knowledge that provides understanding of this world and how it works is called science.

OR

The systematic study of nature based on observation, inference, prediction and experimentation is called science.

(b) Chemistry

The branch of science that deals with the composition, structure, properties and reactions of matter is called chemistry.

It deals with every aspect of life.

Advantages of Chemistry

Following are advantages of development of science and technology.

- **i.** Petrochemical products
- **ii.** Medicines and drugs
- **iii.** Soap and detergents
- **iv.** Paper and plastics
- **v.** Paints and pigments
- **vi.** Insecticides and pesticides
- **vii.** It improves our health and environment
- **viii.** It helps to explore and conserve the natural resources.

Disadvantages of Chemistry

Following are some major disadvantages of chemistry.

- **i.** Generation of toxic waste materials
- **ii.** Contaminated water
- **iii.** Polluted air and contaminated food
- **iv.** Dangerous war weapons

Q.2 Describe the various branches of chemistry.

Ans.

Branches of Chemistry

Chemistry is divided into following main branches:

i. Physical Chemistry

"The branch of chemistry that deals with the relationship between the composition and physical properties of matter along with the changes in them is called physical chemistry."

Scope

Following properties are studied in this branch:

Structure of atoms or formation of molecules, behavior of gases, liquids and solids, the study of the effects of temperature or radiations on matter

ii. Organic Chemistry

"The branch of chemistry that deals with the study of covalent compounds of carbon and hydrogen (hydrocarbons) and their derivatives is called organic chemistry."

Scope

Organic chemists determine the structure and properties of these naturally occurring as well as synthesized compounds.

Scope of this branch covers petroleum, petrochemicals and pharmaceutical industries.

iii. Inorganic Chemistry

"The branch of chemistry that deals with the study of all elements and their compounds except those of compounds of carbon and hydrogen (hydrocarbons) and their derivatives is called inorganic chemistry."

Applications

It has applications in every aspect of the chemical industry such as glass, cement, ceramics and metallurgy (extraction of metals from ores).

iv. Biochemistry

"The branch of chemistry that deals with the study of structure, composition, and chemical reactions of substances found in living organisms is called biochemistry."

Scope

It covers all chemical processes taking place in living organisms such as synthesis and metabolism of bio-molecules like carbohydrates, proteins, fats etc.

Emergence of biochemistry as a separate discipline

Biochemistry emerged as a separate discipline when scientists began to study:

- How living things obtain energy from food
- How' the fundamental biological changes occur during a disease.

Applications:

It is applied in the fields of medicine, food science and agriculture.

v. Industrial Chemistry

"The branch of chemistry that deals with the manufacturing of chemical substances (elements and compounds) on commercial scale, is called industrial chemistry."

Applications:

- It deals with the manufacturing of basic chemicals such as oxygen, chlorine, ammonia, caustic soda, nitric acid and sulphuric acid.
- Use of these chemicals to provide the raw materials for many other industries such as fertilizers, soap, textiles, agricultural products, paints and paper etc

vi. Nuclear Chemistry

"The branch of chemistry that deals with the radioactivity, nuclear processes and properties is called nuclear chemistry."

Scope

Following phenomenon are studied in this branch:

Main concern: The main concern of this branch is with the energy of the atom and its uses in daily life.

The chemical effects resulting from the absorption of radiation within living animals, plants, and other materials are also studied in this branch.

Applications

It has vast applications in medical treatment (radiotherapy), preservation of food and generation of electrical power through nuclear reactors.

vii. Environmental Chemistry

"The branch of chemistry that deals with the study of components of the environment and the effects of human activities on the environment is called environmental chemistry."

Applications

- Environmental chemistry is related to other branches like biology, geology, ecology, soil and water chemistry, mathematics and engineering.
- The knowledge of chemical processes taking place in environment is necessary for its
  improvement and protection against pollution.

viii. Analytical Chemistry

"The branch of chemistry that deals with separation and analysis of a sample to identify its components is called analytical chemistry. The separation is carried out prior to qualitative and quantitative analysis."

Qualitative Analysis

It provides the identity of a substance (composition of chemical species).

Quantitative Analysis

It determines the amount of each component present in the sample.

Scope

- In this branch different techniques and instruments used for analysis are studied.
- The scope of this branch covers food, water, environmental and clinical analyses.

1.2 Basic Definitions

Q.3 Define the following terms.

Matter, Substance, Mixture, Physical properties and Chemical properties

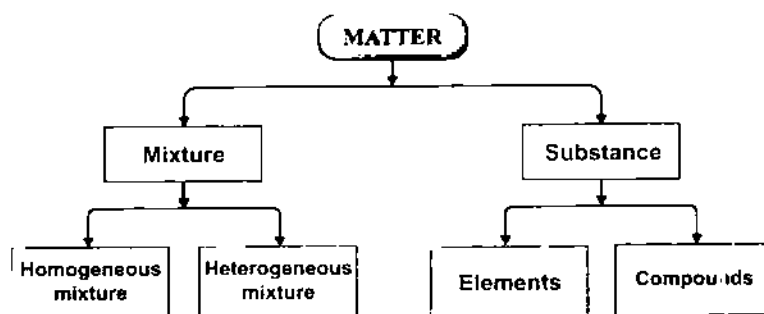
Ans: Matter

"Anything that has mass and occupies space is called matter."

Examples:

Our bodies as well as all the things (air, water, chair etc.) around us are examples of matter.

Physical states of matter:

In chemistry we study all types of matters that can exist in any of three physical states; solid, liquid or gas. All the materials may either be substance or mixture.

Substance

"A piece of matter in pure form is called substance."

Every substance has a fixed composition and specific properties or characteristics

Types of substances

i. Element: iron, gold, silver etc.

ii. Compound: water, carbon dioxide, sulphuric acid etc.

Mixture

"A piece of matter in impure form is called mixture"

Example

Soil, ice cream, milk etc

Types of mixture

i. Homogeneous mixture

ii. Heterogeneous mixture

Physical Properties

"The properties that are associated with the physical state of a matter are called physical properties."

Examples:

Colour, smell, taste, hardness, shape of crystal, solubility, melting or boiling points

Explanation:

When ice is heated, it melts to form water. When water is further heated, it boils to give steam. In this entire process only the physical state of water changes where as its chemical composition remains the same.

iv Chemical Properties

"The properties that depend upon the composition of the substance are called chemical properties"

When a substance undergoes a chemical change, its composition changes and a new substance is formed.

Examples

i. Rusting of iron

ii. Decomposition of water is a chemical change as it produces hydrogen and oxygen gases.

1.2.1 Elements, Compounds and Mixtures

Q.4 Write a detailed note on elements. (Ex. Q.1)

Ans:

Elements

Historical Background:

- **Number of elements in early ages:** In the early ages, only nine elements (carbon, gold, silver, tin, mercury, lead, copper, iron and sulphur) were known.

Previous Definition:

In early ages the element was defined as:

"The substances that could not be broken down into simpler units by ordinary chemical processes are called elements."

- **Number of elements until the end of 19th century:** Until the end of nineteenth century sixty-three elements had been discovered.

Present number of elements:-

- Now 118 elements have been discovered, out of which 92 are naturally occurring elements.

Modern definition

"A substance made up of same type of atoms, having same atomic number and it cannot be decomposed into simple substances by ordinary chemical means is called an element."

Occurrence of elements

Elements occur in nature in free or combined form. All the naturally occurring elements in the world have different percentages in the earth's crust, oceans and atmosphere.

Table 1.1 Natural Occurrences by weight % of some major elements

Crust of Earth		Oceans		Atmosphere	
Oxygen	47%	Oxygen	86%	Nitrogen	78 %
Silicon	28 %	Hydrogen	11%	Oxygen	21%
Aluminum	7.8 %	Chlorine	1.8 %	Argon	0.9%

Physical states of elements:

Elements may be solids, liquids or gases.

i. Solids:

Majority of the elements exist as solids e.g.

Sodium, copper, zinc, gold

ii. Liquid:

There are very few elements which occur in liquid state e.g.

Mercury and bromine

iii. Gases:

A few elements exist as gases e.g

Nitrogen, oxygen, chlorine and hydrogen

Classification of elements:

On the basis of their properties, elements are divided into:

- **Metals:** e.g. sodium, copper. Zinc, gold etc

  About 80 percent of the elements are metals.
- **Non-metals :** e.g.: nitrogen oxygen, carbon etc.
- **Metalloids:** *Elements which have proportion of both metals and non metals are called metalloids or semi matals* e.g. germanium, arsenic, antimony etc.

DO YOU KNOW

- Major part of the body is made up of water i.e. 65% to 80% by mass.
- Six elements constitute about 99% of our body mass; namely: Oxygen 65 %, Carbon 18%, Hydrogen 10 %, Nitrogen 3%, Calcium 1.5% and Phosphorus 1.5%.
- Potassium, Sulphur, Magnesium and Sodium constitute 0.8% of our body mass. Whereas Copper, Zinc, Fluorine, Chlorine, Iron, Cobalt and Manganese constitute only 0.2% of our body mass.

Q.5 Define the symbol. How symbols can be assigned?

Ans: **Symbol:** *"An abbreviation for the name of element is called symbol."*

Derivation of word symbol:

A symbol is taken from the name of that element in English, Latin, Greek or German.

Methods to write down symbol:

- In case of one-letter symbol first capital letter is taken as symbol e.g. H for Hydrogen, N for Nitrogen, and C for Carbon
- In case of two letters symbol, first letter is capital and second well pronounced letter is small e.g. Ca for Calcium, Na for Sodium and C*l* for Chlorine.

Significance of symbol:

i. It represents the name of the element

ii. It indicates one atom of the element.

Q.6 Define valency. Write a detailed note on concept of valency.

Valency

"The combining capacity of an element with other elements is called valency."

Dependence: It depends upon the number of electrons in the outermost shell.

a. Valency of elements in covalent compounds

"In simple covalent compounds it can be defined as:

It is the number of hydrogen atoms which will combine with one atom of that element.

OR

The number of bonds formed by one atom of that element is called valency."

Different Numbers of atoms of hydrogen combine with one atom of these elements to form compounds like.

```
                                  H              H
                                  |              |
                                  N – H      H – C – H
   H – Cl        H – O – H        |              |
                                  H              H
```

The valency of Cl, O, N and C is 1,2,3 and 4 respectively.

b. Valency of elements in ionic compounds

In simple ionic compounds valency is definedas:

"*The number of electrons gained or lost by an atom of an element to complete its octet.*"

i. Elements having less than four electrons in valence shell:-

Elements having less than four electrons in the valence shell prefer to lose the electrons to complete their octet.

Examples:

Na, Mg and Al have 1,2 and 3 electrons in their valence shells respectively and they loose 1,2 and 3 electrons to have valency of 1,2 and 3 respectively.

ii. Elements having more than four electrons in valence shell:

Elements having four or more than four electrons in their valence shell, gain electrons to complete their octet.

Examples:

Nitrogen, Oxygen and Chorine have 5, 6 and 7 electrons in their valence shells respectively. They gain 3, 2 and 1 electrons respectively to complete their octet. Hence they show valency of 3, 2 and I respectively.

Table 1.2 Some Elements and Radicals with their Symbols and Common Valencies

Element / Radical	Symbol	Valency	Element / Radical	Symbol	Valency
Sodium	Na	1+	Hydrogen	H	1+, 1–
Potassium	K	1+	Chlorine	Cl	1–
Silver	Ag	1+	Bromine	Br	1–
Magnesium	Mg	2+	Iodine	I	1–
Calcium	Ca	2+	Oxygen	0	2–
Barium	Ba	2+	Sulphur	S	2–
Zinc	Zn	2+	Nitrogen	N	3–
Copper	Cu	1+, 2+	Phosphorus	P	3–,5+
Mercury	Hg	1+,2+	Boron	B	3+
Iron	Fe	2+, 3+	Arsenic	As	3+
Aluminum	Al	3+	Carbon	C	4+,4–
Chromium	Cr	3+	Carbonate	CO_3^{2-}	2–
Ammonium	NH_4^+	+1	Sulphate	SO_4^{2-}	2–
Hydronium	H_3O^+	1+	Sulphite	SO_3^{2-}	2–
Hydroxide	OH^-	1–	Thiosulphate	$S_2O_3^{2-}$	2–
Cyanide	CN^-	1–	Nitride	N^{3-}	3–
Bisulphate	HSO_4^-	1–	Phosphate	PO_4^{3-}	3–
Bicarbonate	HCO_3^-	1–	Bisulphite	HSO_3^{1-}	1–

Variable Valency

"*Some elements show more than one type of the valency. Valency of such elements is called variable valency.*"

Examples

- In ferrous sulphate ($FeSO_4$), the valency of iron is 2+.
- In ferric sulphate ($Fe_2(SO_4)_3$), the valency of iron is 3+.

Element	Variable Valency	Element	Variable Valency
Copper (Cu)	1 and 2	Lead (Pb)	2 and 4
Iron (Fe)	2 and 3	Phosphorous (P)	3 and 5
Mercury	1 and 2	Sulphur (S)	2,4 and 6
Tin (Sn)	2 and 4		

Note:

Generally, the Latin or Greek name for the element (e.g., Ferrum) is modified to end in 'ous' for the lower valency (Ferrous = 2+) and to end in 'ic' for the higher valency (Ferric = 3+).

Q.7 Describe the compound. How it is classified?

Ans: Compound

Definition:

"Substance made up of two or more elements chemically combined together in a fixed ratio by mass is called compound."

Examples:

Carbon dioxide is a compound by a chemical combination between carbon (C) and Oxygen (O_2)

Water is a compound formed by a chemical combination between hydrogen and oxygen in a fixed ratio of 1:8 by mass.

Properties

i. In compounds, elements lose their own properties and produce new substances (compounds) that have entirely different properties.

ii. Compounds can't be broken down into its constituent elements by simple physical methods.

iii. Elements chemically combine together in a fix ratio by mass and form compound.

iv. All compounds are represented by a simple chemically formula

Classification of compounds

Compounds can be classified as:

i. Ionic compound **ii.** Covalent compound

i. Ionic compounds:

"Compounds that contain oppositely charged ions held together by ionic bonds are called ionic compounds."

Properties:

- Ionic compound do not exist in independent molecular form.
- They form a three-dimensional crystal lattice, in which each ion is surrounded by oppositely charged ions.
- They have high melting and boiling points due to strong attraction between oppositely charged ions.
- These compounds are represented by formula units

Examples: Sodium chloride (NaCl), Potassium bromide (KBr), Copper sulphur

($CuSO_4$), Ferrous sulphate etc

Table 1.3: Some common compounds with their formulae:

Compound	Chemical Formula
Water	H_2O
Sodium chloride (Common salt)	$NaCl$
Silicon dioxide (Sand)	SiO_2
Sodium hydroxide (Caustic Soda)	$NaOH$
Sodium carbonate (Washing Soda)	$Na_2CO_3.10H_2O$
Calcium oxide (Quick Lime)	CaO
Calcium carbonate (Lime Stone)	$CaCO_3$
Sugar	$C_{12}H_{22}O_{11}$
Sulphuric acid	H_2SO_4
Ammonia	NH_3

ii. Covalent compounds

"Compounds formed by the sharing of electrons between different atoms are called covalent compounds."

Examples:

Water (H_2O), Hydrocloric acid (HCl), Sulphuric (H_2SO_4), Methane (CH_4) etc

Properties:

- The Covalent compounds mostly exist in molecular form.
- A molecule is a true representative of the covalent compounds
- They are represented by their molecular formulae

Q.8 What is a mixture? Explain its types.

Ans: Mixture

"A substance made up of two or more elements or compounds physically combined together in any ratio by mass is called mixture."

Properties

- The component substances retain their own chemical identities and properties,
- The mixture can be separated into parent components by physical methods such as distillation, filtration, evaporation, precipitation or magnetization.

Types of Mixtures

Mixtures can be classified as:

i. Homogeneous Mixtures

Mixture that has uniform composition throughout is called homogenous mixture.

Examples: Air, Gasoline, Ice cream etc.

ii. Heterogeneous Mixtures

Mixture that does not have uniform composition throughout is called heterogeneous mixture.

Examples: Soil, rock, wood etc.

DO YOU KNOW

Air is a mixture of nitrogen oxygen, carbon dioxide, noble gases and moisture.
Soil is a mixture of sand, clay, mineral salts, water and air.
Milk is a mixture of calcium, water, sugar, fat, proteins, mineral salts and vitamins
Brass is a mixture of copper and zinc metals

Q.9 What is difference between compound and mixture?

Ans: Difference between a compound and a mixture:

Property	Compound	Mixture
i. Formation	• It is formed by a chemical combination of atoms of elements.	• Mixture is formed by the simple mixing up of the substances.
ii. Properties	• The constituents lose their identity and form a new substance having entirely different properties from them.	• Mixture shows the properties of the constituents.
iii. Ratio	• Compounds always have fixed composition by mass.	• The minimum number and ratio of the components may not be fixed.
iv. Separation of components	• The components cannot be separated by physical means.	• The components can be separated by simple physical methods.
v. Representation	• Every compound is represented by a chemical formula.	• It consists of two or more components and does not have any chemical formula.
vi. Composition	• Compounds have homogeneous composition.	• They may be homogeneous or heterogeneous in composition
vii. Melting point	• A compound has a sharp and fixed melting point.	• A mixture does not have a sharp and fixed melting point.

1.2.1 Atomic Number and Mass Number

Q.10 Define atomic number and mass number. Explain with the help of examples.

Ans:

i. Atomic Number

"The number of protons present in the nucleus of an atom of the element is called atomic number."

Symbol

It is represented by symbol 'Z':

Examples:

All hydrogen atoms have 1 proton, their atomic number Z = 1.

All carbon atoms have 6 protons, their atomic number Z = 6.

All oxygen atoms have 8 protons, their atomic number Z = 8.

All sulphur atoms have 16 protons, their atomic number Z = 16.

In a neutral atom: Atomic number = Number of protons = Number of electrons

Explanation:

As all atoms of an element have the same number of protons in their nuclei, they have the same atomic number. Hence each element has a specific atomic number termed as its identification number.

ii. Mass Number

"The sum of number of protons and neutrons present in the nucleus of an atom is called mass number."

Symbol

It is represented by symbol 'A'.

Examples:

- Hydrogen atom has one proton and zero number of neutron in its nucleus, there fore mass number of hydrogen is

$$A = 1 + 0 = 1$$

- Carbon atom has 6 protons and 6 neutrons, hence its mass number A=12.

Explanation

Mass number = number of protons + number of neutrons

It is calculated as: A=Z+n where n is the number of neutrons. Each proton and neutron has one unit of mass.

Representation

By convention, the mass number is written at the top left corner of the symbol of the atom and atomic number is written at the bottom left corner.

$$^{A}_{Z}X$$

Examples: $^{12}_{6}C$, $^{23}_{11}Na$ etc.

Some elements along with their atomic numbers and mass numbers

Element	Number of Protons	Number of Neutrons	Atomic Number Z	Mass Number A

Hydrogen	1	0	1	1
Carbon	6	6	6	12
Nitrogen	7	7	7	14
Oxygen	8	8	8	16
Fluorine	9	10	9	19
Sodium	11	12	11	23
Magnesium	12	12	12	24
Potassium	19	20	19	39
Calcium	20	20	20	40

1.2.3 Relative atomic mass (A_r) and atomic mass unit

Q.11 Explain the relative atomic mass and atomic mass unit.

Ans: Relative Atomic Mass (A_r)

"The average mass of atoms of an element as compared to $\frac{1}{12}$ (one-twelfth) the mass of one atom of carbon-12 isotope is called relative atomic mass.

$$\text{Relative atomic mass} = \frac{\text{average atomic mass of 1 atom of an element}}{\frac{1}{12} \text{ x mass of 1 atom of carbon - 12}}$$

Examples:

A_r of carbon = 12 amu,

A_r of oxygen = 16 amu

Isotope

"Atoms of an element having different mass number but same atomic number"

Explanation

The mass of an atom is too small to be determined practically. However, certain instruments enable us to determine the ratio of the atomic masses of various elements to that of carbon-12 atoms. This ratio is known as the relative atomic mass of the elements. The standard of relative atomic mass is based on carbon-12 standard; the mass of an atom of carbon is 12 and 1/12th of it comes to be one. When we compare atomic masses of other elements with carbon-12 atoms, they are expressed as relative atomic masses of those elements.

Unit of A_r (Relative atomic mass)

"The unit for relative atomic mass is called atomic mass unit."

Atomic mass unit

"One atomic mass unit is $1/12^{th}$ the mass of one atom of carbon-12.The atomic mass unit is abbreviated as amu."

On atomic mass scale, the atomic mass of carbon-12 is taken as 12.00 amu

$$1 \text{ amu} = \frac{1}{12} \times \text{mass of carbon-12 atom}$$

$$\mathbf{1\ amu} = 1.66 \times 10^{-24} \text{g}$$

Examples:

- Mass of a proton = 1.0073 amu or 1.672×10^{-24}g
- Mass of a neutron = 1.0087 amu or 1.674×10^{-24}g
- Mass of an electron = 5.486×10^{-4} amu or 9.106×10^{-28} g

1.2 How to Write a Chemical Formula

Q.12 (a) Define the chemical formula. Write down the steps to write chemical formula.
(b) What is significance of chemical formula?

Ans: (a) Chemical formula

"Representation of an element or a compound in terms of symbols is called chemical formulae"

Example:

- Chemical formula of aluminium sulphate: $Al_2(SO_4)_3$
- Chemical formula of calcium phosphate: $Ca_3(PO_4)_2$
- Chemical formula of chlorine: Cl_2
- Chemical formula of water: H_2O

Steps to write down chemical formula:

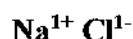
i. Symbols of two elements are written side-by-side, in the order of positive ion first and negative ion later.

ii. The valency of each ion is written on the right top corner of its symbol, e.g. Na^+, Ca^{2+},Cr^{3+} and O^{2-}.

iii. This valency of each ion is brought to the lower right corner of other ion D 'cross-exchange' method e.g Na^+Cl^-, Na^+Cl^-, $NaCl$ $Ca^{+2}Cl^-$, $Ca^{2+}O^{2-}$

$\mathbf{Na^{1+}\ Cl^{1-}}$

$\mathbf{Na_1\ Cl_1}$

NaCl

iv. If the valencies are same, they are offset and are not written in the chemical formula. But if they are different, they are indicated as such at the same position

v. In case of sodium chloride both the valencies are offset and formula is written as NaCl, whereas, calcium chloride is represented by formula $CaCl_2$.

vi. If an ion is a combination of two or more atoms which is called radical, bearing net charge on it. e.g. SO_4^{2-} (sulphate ion) and PO_4^{3-} (phosphate ion) then the net charge represent the valency.

Significance of chemical formula

i. It represents the name of the substance e.g., H_2O (water)

ii. It tells the name of the elements as present in the compound.

iii. It indicates the mass of an element or a compound in amu or grams.

iv. It also represents one mole of the molecule or formula unit in the balanced chemical equation.

v. It is in fact one molecule or formula unit of the substance.

Q.13 Define empirical formula. Describe the empirical formula of ionic and covalent compounds.

Ans: Empirical formula

"A formula that indicates the simplest whole number ratio of atoms of different elements present in a compound is called an empirical formula."

(a) Empirical formula of covalent compounds

"It is simplest whole number ratio of atoms of each element present in a compound"

Examples:

i. Silica or sand (silicon dioxide)	=	Si : O
		1: 2
Thus empirical formula of silica	=	SiO_2
ii. In glucose:	=	C : H : 0
		6 : 12: 6
	=	1: 2: 1
Thus empirical formula of glucose	=	CH_2O

The covalent compound silica (sand) has simplest ratio of 1:2 of silicon and oxygen respectively. Therefore, its empirical formula is SiO_2. Similarly, glucose has simplest ratio 1:2:1 of carbon, hydrogen and oxygen respectively. Hence its empirical formula is CH_2O.

(b) Empirical formula of ionic compounds

The ionic compounds exist in three dimensional network forms. Each ion is surrounded by oppositely charged ions in such a way to electrically neutral compound. Therefore, the simplest unit taken as a represent an of an ionic compound is called formula unit.

Formula unit

"The simplest whole number ratio of ions, as present in the ionic compound is called formula unit. In other words, ionic compounds have only empirical formulae."

Examples:

i. Formula unit of common salt consists of one Na^+ and one Cl^- ion and its empirical formula is NaCl.

ii. Formula unit of potassium bromide is KBr, which is also its empirical formula

1.2.4.2 Molecular Formula

Q.14 What is molecular formula? Write down the relationship between molecular and empirical formula.

Ans: Molecular formula:

"A formula that indicates actual number of atoms of each element present in a molecule of the substance is called molecular formula."

Examples:

i. Molecular formula of water, benzene, chlorine and sulphur are H_2O, C_6H_6, Cl_2 and S_8 respectively.

Derivation of molecular formula: (Relationship between molecular formula and empirical formula)

Molecular formula is derived from empirical formula by the following relationship:

$$\text{Molecular formula} = n \times (\text{Empirical formula})$$

Where

$$n = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$$

Molecular formula of a compound is determined experimentally. The value of 'n' may be 1, 2, 3, 4, 5, 6 & so on.

For example, molecular formula of benzene is C_6H_6 which is derived from the empirical formula CH where the value of n is 6.

Explanation:

i. The molecular formula of a compound may be same or a multiple of the empirical formula.

ii. A few compounds having different empirical and molecular formulae.

iii. Some compounds may have same empirical and molecular formula e.g water (H_2O), hydrochloric acid (HC*l*) etc.

Table: Compounds with their empirical and molecular formulae

Compound	Empirical formula	Molecular formula
Hydrogen peroxide	HO	H_2O_2
Benzene	CH	C_6H_6
Glucose	CH_2O	$C_6H_{12}O_6$

Molecular Mass and Formula Mass

Q.15 Define Molecular mass and formula mass. Give examples.

Ans: i. Molecular Mass

"The sum of atomic masses of all the atoms present in one molecule of a molecular substance is its molecular mass."

Examples:

i. $H_2O = 2(1 \text{ amu}) + 1(16 \text{ amu})$

$= 2 \times 1 + 16$

$= 2 + 16$

$= 18 \text{ amu}$

ii. Molecular mass of carbon oxide (CO_2) is 44 amu

iii. Molecular mass of chlorine (Cl_2) is 71 amu.

ii. Formula Mass

"The sum of atomic masses of all atoms present in one formula unit of an ionic compound is called formula mass."

Some ionic compounds that form three-dimensional solid crystal, are represented by their formula units. Formula mass in such cases is the sum of atomic mass.

Examples:

i. Formula mass of sodium chloride is 5.5 amu

ii. Formula mass of calcium carbonate is 100 amu

1.3 Chemical Species

Q.16 Write a note on the following chemical species.

i. Ions (Cations and anions) **ii.** Molecular ions **iii.** Free radicals

Ans: i. Ions (Cations and Anions)

"An atom or group of atoms having a charge on it is called ion. The charge may be positive or negative."

Types of ions:

There are two types of ions i.e. cations and anions

a. Cations:

"An atom or group of atoms having positive charge on it is called cation."e.g. H^+ , Na^+ , Ca^{2+}

Formation

The cations are formed when an atom loses electrons from its outermost shell.

The following equations show the formation of cations from atoms.

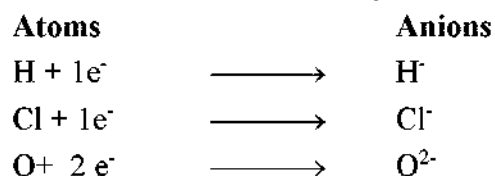
Atoms		Cations
H	$\longrightarrow$	$H^+ + 1e^-$
Na	$\longrightarrow$	$Na^+ + 1e^-$
Ca	$\longrightarrow$	$Ca^{2+} + 2e^-$

b. Anions:

"An atom or a group of atoms that has a negative charge on it, is called anion."e.g. Cl^{1-}, O^{2-}, H^{1-}

Formation

Anion is formed by the gain or addition of electrons to an atom. Following examples show the formation of an anion by addition of electrons to an atom.

Atoms		Anions
$H + 1e^-$	$\longrightarrow$	H^-
$Cl + 1e^-$	$\longrightarrow$	Cl^-
$O + 2e^-$	$\longrightarrow$	O^{2-}

ii. Molecular Ion

"A molecular ion or radical is a species having positive or negative charge on it. When a molecule losses or gains an electron, it forms a molecular ion."

Types of molecular ions:

Molecular ions are of two types.

a. Cationic molecular ion

b. Anionic molecular ion

a. Cationic molecular ion

"The molecular ions which carry positive charge are called cationic molecular ions". They are formed by losing of electron are called cationic molecular ion.

Examples:

N_2^+, He^+, CH_4^+ etc.

b. Anionic molecular ion

"The molecular ions which carry negative charge are called anionic molecular ions."

They are formed by gaining of electron.

Examples:

N_2^-, O_2^{-2}, SO_4^{-2}, PO_4^{-3} etc.

Note: Cationic molecular ions are more abundant than anionic molecular ions.

Generation or formation of molecular ions:

When gases are bombarded with high-energy electrons in a discharge tube, they ionize to give molecular ions

iii. Free Radicals

"Atoms or group of atoms possessing odd number of (unpaired) electrons are called free radicals."

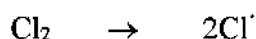
Representation

It is represented by putting a dot over the symbol of an element.

Examples: $Cl^{\bullet}$, $H^{\bullet}$, $OH^{\bullet}$, $CH_3^{\bullet}$ etc.

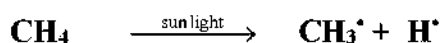
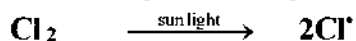
Formation

Free radicals are generated by the hemolytic (equal) breakage of the bond between two atoms when they absorb heat or light energy.

Molecules		Free radicals
Cl_2	→	$2Cl^{\cdot}$
CH_4	→	$H3C^{\cdot} + H^{\cdot}$

Reactivity

A free radical is extremely reactive species as it has the tendency to complete its octet.

$$Cl_2 \xrightarrow{\text{sun light}} 2Cl^{\bullet}$$

$$CH_4 \xrightarrow{\text{sun light}} CH_3^{\bullet} + H^{\bullet}$$

Chemical Species:

"The atoms, molecule ion, molecular ion and free radical are called chemical species."

Example: Na, H_2O, Na^+ Cl^- CH^{+4}, CH_3 and Cl etc.

Simple ion/ monoatomic ion:

"An ion having a single atom is called simple ion"

Examples: Ht, Ct, Mg^{+2}, Al^{+3}, N^{-3} dtc.

Complex/Polyatomic ion:

"An ion having two or more atoms is called a polyatomic ion".

Examples: Ammonium ion aH^{+1}, nitrate ion, NO_3^{1-} etc.

Q.17 What is difference between ions and free radicals?

Ans: Difference between ions and free radicals:

Ions	Free Radicals
i. Definition: *These are the atoms which bear some charge*	**i.** *These are the atoms that have odd number of electrons*
ii. Existence:	**ii.** They can exist in solutions as well in air

They exist in solution or in crystal Lattice	
iii. Effect of light on formation: Their formation is not affected by the presence of light	**iii.** They may form in the presence of light

Q.18 What is difference between molecule and molecular ion?

Ans: Difference between molecule and molecular ion

Molecule	Molecular Ion
i. Definition: *It is the smallest particle of a substance which can exist independently and shows all the properties of that substance (elementor compound)*	**i.** It is formed by gain or loss of electrons by a molecule
ii. Charge: It is always neutral	**ii.** It can have negative or positive charge
iii. Formation: It is formed by the combination of atoms	**iii.** It is formed by the ionization of a molecule
iv. Stability: It is a stable unit	**iv.** It is a reactive specie

Q.19 What is difference between atoms and ions?

Ans: Difference between atoms and ions

Atom	Ion
i. Definition: *Smallest particle of an element which can take part in a chemical reaction and may or may not exist independently is called an atom. It is the smallest particle of an element.*	**i.** *It is the smallest unit of an ionic compound.*
ii. Existence: It can or cannot exist independently and can take part in a chemical reaction.	**ii.** It cannot exist independently and is surrounded by oppositely charged ions.
iii. Charge: It is electrically neutral	**iii.** It has a net charge (either negative or positive) on it.

Q.20 What is meant by molecule? Explain the types of molecules with the help of examples.

Ans: Molecule

A molecule is formed by the chemical combinations of atoms.

Definition:

It is the smallest particle of a substance which can exist independently and shows all the

properties of that substance (element or compound)

Examples: H_2O, Cl_2, S_8, H_2SO_4 etc.

Properties

i. It is the smallest unit of a substance.

ii. It shows all the properties of the substance and can exist independently.

iii. There are different types of molecules depending upon the number and types of atoms combining.

Types of molecules:

a. On the basis of no of atoms

i. Monoatomic Molecule

"A molecule consisting of only one atom is called mono atomic molecule."

Examples:

The inert gases helium, neon and argon all exist independently in atomic form.

ii. Diatomic Molecules

"A molecule consisting of two atoms is called diatomic molecule."

Examples:

Hydrogen (H_2), Oxygen (O_2), Chlorine (Cl_2) and Hydrogen chloride (HCl)

iii. Triatomic molecules:

"A molecule consisting of three atoms is called triatomic molecule."

Examples: H_2O, CO_2 etc.

iv. Polyatomic Molecules

"A molecule consisting of many atoms is called polyatomic molecule."

Examples:

Methane (CH_4), Sulphuric acid (H_2SO_4) and Glucose ($C_6H_{12}O_6$).

b. On the basis of type of atoms:

i. Homoatomic Molecule

"A molecule consisting of same type of atoms is called homoatomic molecule."

Examples:

Hydrogen (H_2), Ozone (O_3), Sulphur (S_8) and Phosphorus (P_4)

ii. Hetroatomic Molecule

"A molecule consisting of different kinds of atoms is called hetroatomic molecule."

Examples: CO_2, H_2O and NH_3

Q.21 Write a note on followings

i. Gram atomic mass **ii. Gram molecular mass** **iii. Gram formula mass**

Ans: Gram Atomic Mass (Gram atom or mole)

"The atomic mass of an element expressed in grams is called gram atomic mass or gram atom. It is also called a mole"

$$\text{No of gram atoms of element} = \frac{\text{Mass of element}}{\text{Atomic mass of element}}$$

Examples:

i. 1 gram atom of hydrogen = 1.008 g = 1 mol of hydrogen

ii. 1 gram atom of carbon = 12.0 g = 1 mol of carbon

It means that 1 gram atom of different elements has different masses.

ii. Gram Molecular Mass

"The molecular mass of an element or a compound expressed in grams is called gram molecular mass. or gram molecule. It is also called a mole."

Examples:

i. 1 gram molecule of water = 18.0 g = 1 mol of water

ii. 1 gram molecule of H_2SO_4 = 98.0 g = 1 mol of sulphuric acid

iii. No of gram molecules of a substance = $\frac{\text{Mass of substance}}{\text{Molecular mass of substances}}$

iii. Gram Formula Mass (Gram formula or mole)

"The formula mass of an ionic compound expressed in grams is called gram formula mass or gram formula. This is also called a mole."

Examples:

i. 1 gram formula of NaCl = 58.5 g = 1 mol of sodium chloride

ii. 1 gram formula of $CaCO_3$ = 100 g = 1 mol of calcium carbonate

1.5 Avogadro's Number and Mole

Q.22 Explain the Avogadro's number.

Ans: Avagadro's Number

Definition

"The number of particles in one mole of a substance is called avagodro number."

Representation and numerical value:

Avogadro's number is represented by symbol 'N_A'. Its numerical value is 6.02×10^{23}.

Introduction:

It is a huge number. It was suggested by an Italian scientist Amaedo Avagadro.

Explanation

In simple words 6.02×10^{23} particles are equal to one mole as twelve eggs are equal to one dozen. To understand the relationship between the Avogadro's number and the mole of a substance lets see the following examples.

Examples:

i. Gram atomic mass of C = 12g = 1 mole of C = 6.02×10^{23} atoms of C

ii. Gram molecular mass of H_2O = 18g = 1 mole of H_2O = 6.02×10^{23} molecules of H_2O

iii. Gram formula mass of NaCl = 58.5g = 1mole of NaCl= 6.02×10^{23} formula units of NaCl

Importance:

In chemistry we deal with substances which are composed of atoms, molecules or formula units. The counting of these particles is not possible for the chemists. The concept of Avogadro's number facilitated the counting of particles contained in the given mass of a substance. Avogadro's number is a collection of 6.02×10^{23} particles.

Explanation with further examples:

For further explanation about number of atoms in molecular compounds or number of

ions in ionic compounds, let us discuss two examples:

i. One molecule of water is made up of 2 atoms of hydrogen and 1 atom of oxygen, hence $2 \times 6.02 \times 10^{23}$ atoms of hydrogen and 6.02×10^{23} atoms of oxygen constitute one mole of water.

ii. One formula unit of sodium chloride consists of one sodium ion and one chloride ion. So there are 6.02×10^{23} number of Na^+ ions and 6.02×10^{23} Cl^- ions in one mole of sodium chloride. Thus, the total number of ions in 1 mole of NaCl is 12.04×10^{23} or 1.204×10^{24}.

Q.23 Define the mole. How mole is helpful for the calculation of particles.

Ans: Mole (chemist secret unit)

Quantitative definition:

"A mole is defined as the amount (mass) of a substance that contains 6.02×10^{23} number of particles (atoms, molecules or formula units) is called a mole."

Qualitative definition:

"The atomic mass molecular mass formula mass or ionic mass of a substance expressed in grams is called a mole."

Symbol:

It is abbreviated as "mol" when it is used as a unit.

Explanation:

Mass of a substance is either one of the following: atomic mass, molecular mass or formula mass. These masses are expressed in atomic mass units (amu). But when these masses are expressed in grams, they are called as molar masses or molar mass of a substance.

Quantitative definition of mole

It is the atomic mass, molecular mass or formula mass of a substance expressed in grams is called mole.

Examples:

i. Atomic mass of carbon expressed in gram = 12g = 1 mole of carbon

ii. Molecular mass of H_2O expressed in gram =18g = I mole of water

iii. Molecular mass of H_2SO_4 expressed in gram = 98 g = 1 mole of H_2SO_4

iv. Formula mass of NaCl expressed in gram = 58.5g = 1 mole of NaCl

Relationship between mole and mass:

$$\text{Number of moles} = \frac{\text{known mass of substance}}{\text{molar mass of substance}}$$

Mass of substance (g) = number of moles × molar mass

Relationship between mole and number of particles:

Number of particles = number of moles × 6.02×10^{23}

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 1: Fundamentals of Chemistry
Extra MCQ's

1. **The branch of chemistry which deals with the study of all elements and their compounds except compounds of carbon and hydrogen and their derivatives is called.**
(a) Organic chemistry (b) Physical chemistry
(c) Inorganic chemistry (d) Bio chemistry
2. **Which one of the followings provides the identity of a substance?**
(a) Qualitative analysis (b) Clinical analysis
(c) Quantitative analysis (d) Chemical analysis
3. **Which one of the followings is applicable in chemical industry like metallurgy, ceramics and glass?**
(a) Organic chemistry (b) Inorganic chemistry
(c) Industrial chemistry (d) Nuclear chemistry
4. **Anything that has mass and occupies space is called:**
(a) Substance (b) Matter (c) Element (d) Atomic mass
5. **The number of elements known in early ages is:**
(a) 118 (b) 109 (c) 63 (d) 9
6. **Until the end of 19^{th} century how many elements were discovered?**
(a) 9 (b) 63 (c) 92 (d) 118
7. **The percentage of aluminum in crust of earth is:**
(a) 47% (b) 28% (c) 7.8% (d) 1.8%
8. **Which one of the following elements is liquid at room temperature?**
(a) Bromine (b) Mercury (c) Nitrogen (d) Both a & b
9. **Piece of matter in pure form is called:**
(a) Mixture (b) Matter (c) Substance (d) Compound
10. **The quantity of potassium, magnesium, sulphur and sodium in human body is:**
(a) 0.2% (b) 0.6% (c) 0.8% (d) 0.4%
11. **The unique property of an element is called:**
(a) Electronegativity (b) Valency (c) Mobility (d) Oxidation number
12. **In water fixed ratio of hydrogen and oxygen by mass is:**
(a) 8:1 (b) 2:16 (c) 1:8 (d) 2:1
13. **The mixture which has uniform composition throughout is called:**
(a) Simple mixture (b) homogeneous mixture
(c) Heterogeneous mixture (d) Compound mixture
14. **Which one of the followings is the mixture of oxygen, nitrogen and carbon dioxide?**
(a) Soil (b) Water (c) Air (d) Brass
15. **The sum of protons and neutrons in the nucleus of an atom is called:**
(a) Atomic number (b) Mass number (c) Formula mass (d) Atomic mass unit
16. **A substance whose atoms have the same atomic number is called:**
(a) Element (b) Substance (c) Mixture (d) Compound

17. **H_2O is an example of:**
(a) Radical cation (b) Radical anion (c) Compound (d) Element

18. **One mole of any material contains chemical units equal to:**
(a) 6.02×10^{23} (b) 6.02×10^{24} (c) 6.02×10^{-24} (d) 6.02×10^{25}

19. **Ion, molecular ions, free radicals and neutral molecules are called:**
(a) Ionic species (b) Molecular species (c) Chemical species (d) Atomic species

20. **Noble gases are the best examples of:**
(a) Mono atomic molecules (b) Hetero atomic molecules
(c) Poly atomic molecules (d) Diatomic molecules

21. **Brass is the mixture of:**
(a) Cu+Zn (b) Cu+Pb (c) Zn+Pb (d) Sn+Cu

22. **The molar mass of H_3PO_4 is:**
(a) 58.5g (b) 98g (c) 40g (d) 98amu

23. **The number of particles in one mole of a substance is called:**
(a) Atomic number (b) Particle number (c) Avogadro's number (d) Mass number

24. **1amu (atomic mass unit) is equivalent to:**
(a) 1.66×10^{23}g (b) 1.66×10^{-24}g (c) 1.66×10^{-23}g (d) 6.02×10^{-25}

25. **The sum of atomic masses of all the atoms in one formula unit of a substance is called:**
(a) Atomic mass (b) Mass number (c) Formula mass (d) Atomic mass unit

26. **Which one of the followings shows the simplest whole number ratio of atoms in a substance?**
(a) Molecular formula (b) Empirical formula
(c) Chemical formula (d) Covalent formula

27. **Chemical formula of washing soda is:**
(a) $Na_2CO_3.H_2O$ (b) $Na_2CO_3.10H_2O$ (c) $Na_2CO_3.7H_2O$ (d) Na_2CO_3

28. **Mass of an electron is:**
(a) 5.486×10^{-4}amu (b) 9.106×10^{-28}amu (c) 1.672×10^{-24} g (d) 1.677×10^{-24} g

29. **Which one of the followings is a reactive species?**
(a) Molecule (b) Molecular ion (c) Compound (d) Formula unit

30. **An atom or group of atoms having odd number of electrons is called:**
(a) Radical (b) Ion (c) Free radical (d) Molecular ion

31. **Which one of the followings is tri atomic molecule?**
(a) H_2SO_4 (b) N_2 (c) CO_2 (d) HCl

32. **The formula mass of an ionic compound expressed in grams is called**
(a) Gram formula mass (b) Gram formula
(c) Mole (d) All of these

33. **The total number of ions in one mole of NaCl is:**
(a) 12.04×10^{23} ions (b) 1.204×10^{23} ions
(c) 6.04×10^{23} ions (d) 61.04×10^{23} ions

34. **Number of moles$\times6.02\times10^{23}$ is equal to:**
(a) Number of moles (b) Number of particles
(c) Mass of substance (d) Mass of particles

35. **The symbol for nitride radical is:**
(a) CO_3^{2} (b) N^{3-} (c) PO_4^{3-} (d) $S_2O_3^{2-}$

36. **It is a liquid element at room temperature**
(a) Mercury (b) Nickel (c) Hydrogen (d) Iodine

37. **Mass of proton is equivalent to:**
(a) 1.672×10^{-24}amu (b) 16.72×10^{-24}g (c) 1.672×10^{24}g (d) 1.672×10^{-24}g

38. The number of neutrons in $_{92}U^{238}$ is:
(a) 194 (b) 92 (c) 146 (d) 238

39. An atom or group of atoms having positive charge on it is called:
(a) Cation (b) Anion (c) Molecule (d) Atom

40. Which of the followings shows variable valency?
(a) Ca (b) Fe (c) B (d) I

41. Which one of the following elements is found most abundantly in the earth crust?
(a) Oxygen (b) Aluminium (c) Silicon (d) Iron

42. The third abundant gas found in the atmosphere is:
(a) Carbon monoxide (b) Oxygen (c) Nitrogen (d) Argon

43. Total number of elements which have been discovered till now are:
(a) 110 (b) 115 (c) 118 (d) 102

44. Which element is found in liquid state?
(a) Gold (b) Bromine (c) Lead (d) Silver

45. The compound which has same empirical formula is:
(a) C_6H_6 (b) $C_6H_{12}O_6$ (c) CH_3COOH (d) Both b and c

46. CaO is the chemical formula of:
(a) Lime stone (b) Quick lime (c) Caustic soda (d) Slaked lime

47. How many elements occur naturally?
(a) 92 (b) 96 (c) 98 (d) 100

48. HSO_4^{1} is the symbol of which one of the following?
(a) Ammonium ion (b) Cyanide (c) Bisulphate (d) Bicarbonate

49. The formula mass of K_2SO_4 is:
(a) 174 amu (b) 174g (c) 170 amu (d) 170 g

50. Which one of the followings has sharp and fixed melting point?
(a) Compound (b) Mixture (c) Both (d) None of these

ANSWER KEY

1	c	11	b	21	a	31	c	41	a
2	a	12	c	22	b	32	d	42	d
3	b	13	b	23	c	33	a	43	c
4	b	14	c	24	b	34	b	44	b
5	d	15	b	25	c	35	b	45	d
6	b	16	a	26	b	36	a	46	b
7	b	17	c	27	b	37	d	47	a
8	d	18	a	28	a	38	c	48	c
9	c	19	c	29	b	39	a	49	a
10	c	20	a	30	c	40	b	50	a

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 2: Structure of Atoms
Exercise Questions

Exercise Multiple Choice Question Answers

1. **Which one of the following results in the discovery of proton**
(a) cathode rays (b) canal rays (c) x-rays (d) alpha rays.
2. **Which one of the following is the most penetrating?**
(a) protons (b) electrons. (c) neutrons (d) alpha particles
3. **The concept of orbit was used by**
(a) J. J. Thomson (b) Rutherford (c) Bohr (d) Planck
4. **Which one of the following shell consists of three subshells.**
(a) O-shell (b) N shell (c) L shell (d) M shell
5. **Which radioisotope is used for the diagnosis of tumor in the body?**
(a) cobalt-60 (b) iodine-131 (c) strontium-90 (d) phosphorus-32
6. **When U-235 breaks up, it produces:**
(a) electrons (b) neutrons (c) protons (d) nothing
7. **The p subshell has:**
(a) one orbital (b) two orbitals (c) three orbitals (d) four orbitals
8. **Deuterium is used to make:**
(a) light water (b) heavy water (c) soft water (d) hard water
9. **The isotope C-12 is present in abundance of:**
(a) 96.9 % (b) 97.6 % (c) 99.7 % (d) none of these
10. **Who discovered the proton:**
(a) Goldstein (b) J. T. Thomson (c) Neil Bohr (d) Rutherford

ANSWR KEY

1	b	3	c	5	a	7	c	9	d
2	c	4	d	6	b	8	b	10	a

Exercise Short Question Answers

Q.1 What is the nature of charge on cathode rays?

Ans: Cathode rays are negatively charged particles. J.J. Thomson discovered the e/m (charge/mass) ratio of cathode rays and found it equal to electron.

Q.2 Give five characteristics of cathode rays.

Ans: The characteristics of cathode rays are as under:

i. These rays travel in a straight line perpendicular to the cathode surface.
ii. They raise the temperature of the body on which they fall.
iii. Light is produced when these rays hit the sides of discharge tube.
iv. They can cast a sharp shadow of an opaque object if placed in their path.
v. The nature of rays does not depend upon the nature of as used in discharge tube.

Q.3 The atomic symbol of a phosphorus ion is given as $\left({}^{31}_{15}P^{3-}\right)$

(a) How many protons, electrons and neutrons are there in the ion?
(b) What is name of the ion?
(c) Draw the electronic configuration of the ion.
(d) Name the noble gas which has the same electronic configuration as the phosphorus ion has.

Ans:

a. In $\left[{}^{31}_{15}P^{3-}\right]$ ion:

i. Number of protons = 15
ii. Number of electron = 15+3=18 (P^{3-} has three more electrons than neutral P-atom)
iii. Number of neutrons = 31-15=16

b. The name of ion is Phosphide ion

c. Electronic configuration of ${}^{31}_{15}P^{3-}$ = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ (P^{3-} has three more electrons than neutral P-atom)

d. Argon has same electronic configuration as the phosphorous ion has.

Q.4 Differentiate between shell and subshell with examples of each.

Ans:

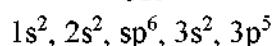
Shell	Sub-shell
i. The circular path of an electron around the nucleus is called shell or principal energy level is called a shell.	*i. Each shell consists of smaller paths called subshells.*
ii. The shells are subdivided into subshells.	**ii.** The subshells are further composed of atomic orbitals.
iii. These are represented by K, L, M, N etc.	**iii.** Example: s, p, d and f are considered as the subshells of a shell. These are represented by s, p, d, f.

Q.5 An element has an atomic number 17. How many electrons are present in K, L and M shells of the atom?

Ans: Atomic number of element = number of electrons = 17
Therefore, its electronic configuration will be

K	L	M
2	8	7

OR

$1s^2, 2s^2, sp^6, 3s^2, 3p^5$

Q.6 Write down the electronic configuration of Al^{3+}. How many electrons are' present in its outermost shell?

Ans: Atomic number of Al = 13

Number of electrons of Al = 13

Number of electrons of Al^{3+} = 13-3 = 10 electrons.

Thus electronic configuration of Al^{3+} ion

K L

2 8

In terms of subshell: $1s^2, 2s^2, sp^6$

Therefore,

Number of electrons present in outer most shell of Al^{3+} = 8 electrons

Q.7 Magnesium has electronic configuration 2, 8, 2,

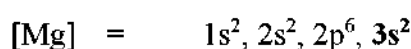
(a) How many electrons are in the outermost shell?

(b) In which subshell of the outermost shell electrons are present?

(c) Why magnesium tend to lose electrons?

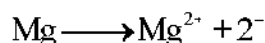
Ans:

a. Electronic configuration of Mg = K L M / 2 8 2

[Mg] = $1s^2, 2s^2, 2p^6, \mathbf{3s^2}$

It has two electrons in the outermost shell.

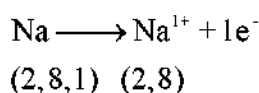
b. The outermost electrons are present in "s" subshell of the 3rd shell (M).

c. Magnesium is electropositive in character. It has the ability to lose its two electrons from its outermost shell.

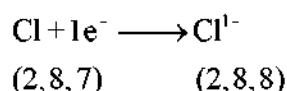
$$Mg \longrightarrow Mg^{2+} + 2^-$$

Q.8 What will be the nature of charge on an atom when it loses an electron or when it gains an electron?

Ans: When an atom loses an electron, it acquires positive charge due to more number of protons in the nucleus e.g.

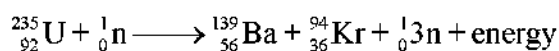
$$Na \longrightarrow Na^{1+} + 1e^-$$

(2,8,1) (2,8)

When an atom gains an electron, it possesses negative charge due to more electrons

than protons in the atom e.g.

$$Cl + 1e^- \longrightarrow Cl^{1-}$$
$$(2,8,7) \qquad (2,8,8)$$

Q.9 For what purpose is U-235 used?

Ans: Radioactive isotope U-235 is used to generate electricity.

$${}^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{139}_{56}Ba + {}^{94}_{36}Kr + {}^{1}_{0}3n + \text{energy}$$

Q.10 A patient has goiter, how will it be detected?

Ans: Isotopes of iodine-131 are used for diagnosis of goiter in thyroid gland. These radioactive isotopes are used as tracers in medicine to diagnose the presence of tumor in the human body.

Q.11 Give three properties of positive rays.

Ans: Positive rays are also called "canal rays". Their properties are:

i. These rays travel in straight line in a direction opposite to the cathode rays.

ii. These are positively charged rays.

iii. Mass of these particles was found equal to that of a proton or simple multiple of it.

Q.12 What are the defects of Rutherford's atomic model?

Ans: Rutherford's atomic model had following defects:

i. Stability of atom:

According to classical theory of radiations, electrons being charged particles should release or emit energy continuously. They should ultimately fall into the nucleus.

ii. Nature of spectrum:

If the electrons emit energy continuously they should form a continuous atomic spectrum but in fact, line atomic spectrum was observed.

Q.13 As long as electron remains in an orbit, it does not emit or absorb energy. When does it emit or absorb energy?

Ans: Electrons do not emit or absorb energy till they remain in their orbits. Electron emits energy when it jumps from high energy level to the lower energy level. An electron absorbs energy when it jumps from a lower energy orbit to a higher energy orbit. The change in energy is given by the following Planck's equation

$$AE = E_2 - E = hv \qquad \text{(Energy absorb)}$$

Where

E_1 = energy of lower energy orbit

E_2 = energy of higher energy orbit

"h" is Planck's" constant. Its value is 6.63×10^{-34} Js and frequency of light.

And

$E_2 - E_1 = -hv$ (Energy emitted)

Exercise Long Question Answers

Q.1 How are cathode rays produced? What are its five major characteristics?

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Q.2 How was it proved that electrons are fundamental particles of an atom?

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Q.3 Draw a labeled diagram to show the presence of protons in the discharge tube and explain how canal rays were produced.

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Q.4 How Rutherford discovered that atom has a nucleus located at the centre of the atom?

Ans: See Q. No. **5** (Subjective Part, Long Questions Answers)

Q.5 One of the postulates of Bohr's atomic model is that angular momentum of a moving electron is quantized. Explain its meaning and calculate the angular momentum of third orbit (i.e. n=3)

Ans:

Q.6 How did Bohr prove that an atom must exist?

Ans: See Q. No. **6** (Subjective Part, Long Questions Answers)

Q.7 What do you mean by electronic configuration? What are basic requirements while writing electronic configuration of an element (atom)?

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.8 Describe the electronic configuration of Na^+, Mg^{2+} and Al^{3+} ions. Do they have the same number of electrons in the outermost shell?

Ans:

(i) Na^+

Electronic configuration is shells = $\begin{matrix} K & L \\ 2 & 8 \end{matrix}$

In terms of subshell : $is^2, 2s^2, sp^6$

(ii) Mg2+:

Electronic configuration in shell: $\begin{matrix} K & L \\ 2 & 8 \end{matrix}$

In terms of subshell: $1s^2, 2s^2, sp^6$

(iii) Al3+

Electronic configuration in shell: K L / 2 8

In terms of subshell: $1s^2$, $2s^2$, sp^6

Hence:

It is proved that all have 8 electrons in their outermost shells.

Q.9 Give the applications of isotopes in the field of radiotherapy and medicines.

Ans: See Q. No. **13** (Subjective Part, Long Questions Answers)

Q.10 What is an isotope? Describe the isotopes of hydrogen with diagrams.

Ans: See Q. No. **11** (Subjective Part, Long Questions Answers)

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 2: Structure of Atoms
Long Questions

Theories and experiments related to structure of atom

Q.1 What is the Dalton's atomic theory and Plum pudding theory?

Ans: Dalton's atomic theory

In the beginning of 19th century John Dalton put forward Atomic Theory. According to it 'all matter is made up of very small indivisible particles called atoms'.

i. All matter is composed of atoms.

ii. An atom is an indivisible, hard, dense sphere.

iii. Atoms of the same element are alike.

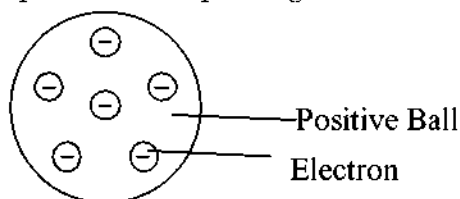
iv. They combine in different ways to form compounds.

J.J Thomson (1856-1940) was a British physicist. He was awarded the 1906 Noble Prize in Physics for the discovery of electron and for his work on the conduction of electricity in gases.

In the light of Dalton's atomic theory, scientists performed a series of experiments. But in the late 1800's and early 1900's, scientists observed new sub-atomic particles.

Plum pudding theory

In the late 1800, J.J Thomson proposed a model of an atom based on coulomb's law. Thomson put forth his "plum pudding" theory.

"He postulated that atoms were solid structures of positively charge with tiny negative particles stuck inside. It is like plums in the pudding."

Q.2 How the cathode rays were discovered? What are its major properties? (Ex. Q.2)

Ans: Cathode rays and discovery of electron: (Ex. Q.1)

Sir William Crooks (1832-1919) was a British chemist and physicist. He was pioneer of a vacuum tubes. He worked on spectroscopy.

Introduction

In 1879 Sir William Crooks performed experiments by passing electric current through gases in a discharge tube at very low pressure.

Experiment:

He took a glass tube fitted with two metallic electrodes, which were connected to a high voltage battery. The pressure inside the tube was kept 10^{-4} atm. When high voltage current was passed through the gas, shiny rays were emitted from the cathode surface move towards the anode as shown in figure 2.1.

Name Reasons:

These rays were given the name of 'cathode rays' as these were originated from the cathode.

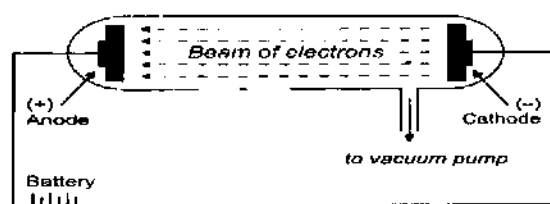



Fig 2.1 Discharge tube used for the production of cathode rays.

Properties of cathode rays:

The major characteristics of cathode rays are given below:

i. **Travelling in straight line:** These rays travel in a straight line perpendicular to the cathode surface.
ii. **Costing of shadow:** They can cast a sharp shadow of an opaque object if placed in their path.
iii. **Deflection in electric field:** They are deflected towards positive plate in an electric field showing that they are negatively charged.
iv. **Rise in temperature:** They raise temperature of the body on which they fall.
v. **e/m ratio:** J.J. Thomson discovered their charge/mass (e / m) 'ratio:
vi. **Production of light:** Light is produced when these rays hit the sides of the discharge tube.
vii. **Nature of cathode:** It was found that the same type of rays were emitted, no matter which gas and which cathode was used in the discharge tube.

Conclusions:

i. All these properties suggested that the nature of cathode rays was independent of the nature of the gas present in the discharge tube or material of the cathode.
The fact that they cast the shadow of an opaque object suggested that these are not rays but they are fast moving material particles. They were given the name electrons.
ii. Since all the materials produce same type of particles, it means all the materials contain electrons.
iii. As we know materials are composed of atoms, hence the electrons are fundamental particles of atoms.

Q.3 How the protons were discovered? Write down its properties. (Ex. Q.3)

OR

Draw a labeled diagram to show the presence of protons in the discharge tube and explain how canal rays were produced.

Ans: Discovery of proton

Introduction:

Protons were discovered by Goldstein in 1886.

Experiment:

Goldstein observed that in addition to cathode rays, other rays were also present in the discharge tube. These rays were traveling in opposite direction to cathode rays. He used a discharge tube having perforated cathode as shown in figure 2.2. He found that these rays passed through holes present in the cathode and produced a glow on the wall. He called these rays as "canal rays".

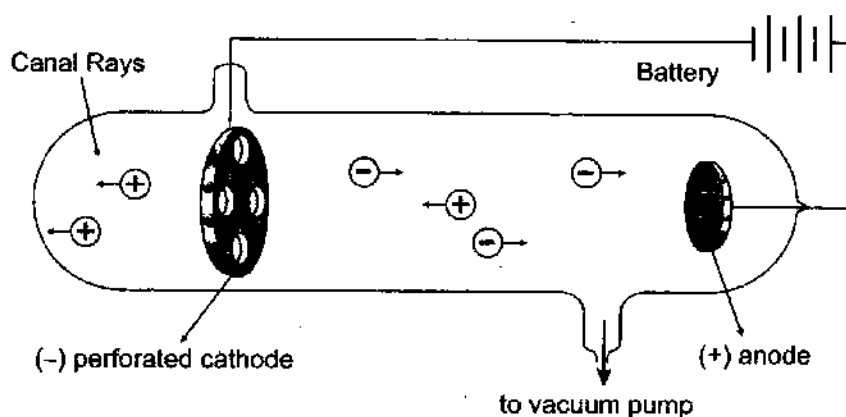



Fig 2.2 Discharge tube used for the production of canal rays.

Properties of positive rays or canal rays:

i. **Travelling in straight line:** These rays travel in a straight line in a direction opposite to cathode rays.

ii. **Deflection in electric field:** Their deflection in electric and magnetic field proved that these were positively charged.

iii. **Dependence:** The nature of canal rays depends upon the nature of gas, present in the discharge tube.

iv. **Origin:** These rays do not originate from the anode. In fact these rays are produced when the cathode rays or electrons collide with the residual gas molecule present in the discharge tube and ionize them.

v. **Mass of positive rays:** Mass of these particles was found equal to that of a proton or simple multiple of it. The mass of a proton is 1840 times more than that of an electron.

Results:

i. These rays are made up of positively charged particles.

ii. The mass and charge of these particles depend upon the nature of the gas in the discharge tube. Hence, different gases produce different types of positive rays having particles of different masses and different charges.

iii. Positive particles produced by a gas will be of the same type i.e. positive rays produced by the lightest gas hydrogen contain protons.

Q.4 How were neutrons discovered? Write down their properties.

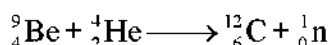
Ans: Discovery of Neutron

Historical Backgrounds:

Rutherford observed that atomic mass of the element could not be explained on the basis of the masses of electron and proton only. He predicted in 1920 that some neutral particle having mass equal to that of proton must be present in an atom. Thus scientists were in search of such a neutral particle.

Experiment:

In 1932 Chadwick discovered neutron, when he bombarded alpha particles on a beryllium target. He observed that highly penetrating radiations were produced. These radiations were called neutron.

$$^{9}_{4}Be + ^{4}_{2}He \longrightarrow ^{12}_{6}C + ^{1}_{0}n$$

Properties

i. **Charge:** Neutrons carry no charge i.e. they are neutral
ii. **Penetration:** They are highly penetrating.
iii. **Mass:** Mass of these particles was nearly equal to the mass of a proton.

Q.5 How Rutherford discovered that atom has a nucleus located at the centre of the atom? (Ex. Q.4)

OR

Explain the Rutherford's atomic structure experiment and atomic model in detail.

Ans: Rutherford's Experiment

(Gold Foil Experiment / α-Scattering Experiment / Atomic Structure Experiment)

Introduction:

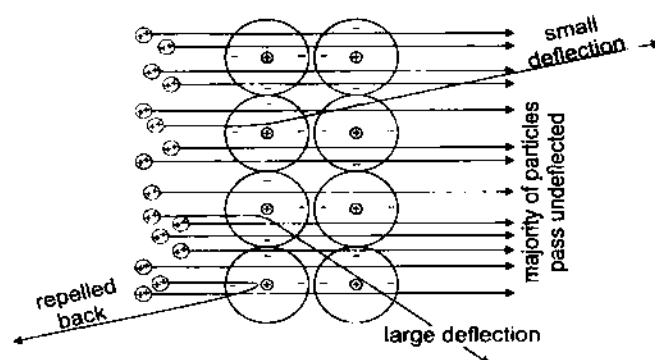
This experiment was performed by Lord Rutherford and his co-worker in 1911. For his work he was awarded |Nobel prize for chemistry in 1908.

Objectives:

Rutherford performed 'Gold Foil' experiment to understand how negative and positive charges could coexist in an atom.

Experiment:

He bombarded alpha particles on a 0.00004 cm thick gold foil. Alpha particles are emitted by radioactive elements like radium and polonium: These are actually helium nuclei (He^{2+}). They can penetrate through matter to some extent.

He observed the effects of a-particles on a photographic plate or a screen coated with zinc sulphide He proved that the 'plum-pudding' model of the atom was not correct.

Rutherford was a British New Zealand chemist. He performed a series of experiments using a particles. He won the 1908 Noble Prize in Chemistry. In 1911, he proposed the nuclear model of the atom and performed the first experiment to split atom. Because of his great contributions, he is considered the father of nuclear science.

Scattering of alpha particles by the atoms of gold foil

Observations:

The observation made by Rutherford were as follows:

i. Almost all the particles passed through the foil un-deflected.

ii. Out of 20000 particles, only a few were deflected at fairly large angles and very few bounced back on hitting the gold foil.

Results of the experiment (Postulates of Rutherford's Atomic Model:

Rutherford proposed planetary model for an atom and concluded following results:

i. Empty part: since most of the particles passed through the foil un-deflected, therefore most of the volume occupied by an atom is empty.

ii. Center of positive charges: The deflection of a few particles proved that there is a 'center of positive charges' in an atom, which is called 'nucleus' of an atom.

iii. Dense and hard nucleus: The complete rebounce of a few particles show that the nucleus is very dense and hard.

iv. Size of nucleus: Since a few particles were deflected it shows that the size of the nucleus is very small as compared to the volume of an atom.

v. Revolving of electron: The electrons revolve around the nucleus.

vi. Number of electrons and protons: An atom as a whole is neutral, therefore the number of electrons in an atom is equal to the number of protons.

vii. Nucleon number: Except electrons, all other fundamental particles that lie within a nucleus are known as nucleons.

Defects in Rutherford's Model:

Although Rutherford's experiment proved that the plum pudding model of an atom was not correct, yet it had following defects:

i. Stability of atom: According to classical theory, electrons being the charged particles should release or emit energy continuously and they should ultimately fall into the nucleus.

ii. Nature of spectrum: If the electrons emit energy continuously, they should form a continuous spectrum but in fact, line spectrum was observed.

Despite of objections on the Rutherford's atomic model, yet it cultivated thought provoking ideas among them.

Q.6 Write down the postulates of Bohr's atomic theory.

OR

How Bohr prove that an atom must exist?

Ans: Bohr's atomic theory:

Introduction

Neil Bohr presented another model of atom in 1913, keeping in view the defects in Rutherford's atomic model.

Basis of Bohr's Atomic Theory

The Quantum Theory of Max Planck was used as foundation for this model.

Energy of an electron:

According to Bohr's model revolving electron in an atom does not absorb or emit energy continuously. The energy of a revolving electron is 'quantized' as it revolves only in orbits of fixed energy, called 'energy levels' by him.

Niels Bohr was a Danish physicist who joined Rutherford in 1912 for his post doctoral research. In 1913, Bohr presented his atomic model based upon Quantum theory. He won the 1922 Noble Prize for Physics for his work on the structure of an atom.

Postulates of Bohr's atomic theory:

i. **Structure of hydrogen atom:** The hydrogen atom consists of a tiny nucleus and electrons are revolving in one of circular orbits of radius or' around the nucleus.

ii. **Energy of orbit:** Each orbit has a fixed energy that is quantized.

iii. **Emission or absorption of energy:** As long as electron remains in a particular orbit it does not radiate or absorb energy. The energy is emitted or absorbed only when an electron jumps from one orbit to another.

iv. **Change in energy:** When an electron jumps from lower orbit to higher orbit it absorbs energy and when it jumps back from higher orbit to lower orbit it radiates energy. This change in energy, ΔE is given by following Planck's equation

Fig 2.4 Bohr's atomic model showing orbits.

$$\Delta E = E_2 - E_1 = h\upsilon$$

Where, h is Planck's constant equal to 6.63×10^{-34} J s, and v is frequency of light.

v. **Angular momentum:** Electron can revolve only in orbits of a fixed angular moment mvr, given as:

$$mvr = n\frac{h}{2\pi}$$

Where 'n' is the quantum number or orbit number having values 1, 2, 3 and so on.

Q.7 What are differences between Rutherford's Atomic theory and Bohr's Atomic theory?

Ans: Differences between "Rutherford's and Bohr's Atomic Theories"

	Rutherford's Atomic Theory	Bohr's Atomic Theory
i.	It was based on classical theory.	It was based upon quantum theory.
ii.	Electrons revolve around the nucleus.	Electrons revolve around the nucleus in orbits of fixed energy.
iii.	No idea about orbits was introduced.	Orbits had angular momentum.
iv.	Atoms should produce continuous spectrum.	Atoms should produce line spectrum.
v.	Atoms should collapse.	Atoms should exist.

Q.8 Write a note on shells and subshells.

Ans: a. Shell

The circular path of an electron around the nucleus is called shell or principal energy level.

Examples: K, L, M, N etc.

Properties of a shell:

- **i.** Shells are the main energy levels that electrons occupy.
- **ii.** Shells are represented by circles around the nucleus.
- **iii.** The number of electron that a shell can accommodate is given by $2n^2$, where 'n' is the shell number.
- **iv.** Different energy levels or shells are counted from the centre to outwards.
- **v.** A shell also consists of subshells or orbitals.
- **vi.** Each subshell or orbital is designated by a small alphabetical letter s, p, d etc.
- **vii.** Energy levels are represented by on' values 1, 2, 3 and so on.
- **viii.** Shells are designated by the alphabets or shells K, L, M and so on.
- **ix.** A shell closer to the nucleus is of minimum energy.
- **x.** Since K shell is closest to the nucleus, the energy of shells increases from K shell and onwards.

Shells and their energies:

1^{st} energy level is K shell; it has the lowest energy.

2^{nd} energy level is L shell; it has more energy than K shell.

3^{rd} energy level is M shell; it has more energy than K and L shell.

4^{th} energy level is N shell; it has more energy than K, L and M shell

Maximum capacity of shells to accommodate electrons:

The number of electron that a shell can accommodate is given by $2n^2$, where 'n' is the shell number.

K shell: $2n^2 = 2(1)^2 = 2\times1 = 2$

L shell: $2n^2 = 2(2)^2 = 2\times4 = 8$

M shell: $2n^2 = 2(3)^2 = 2\times9 = 18$

N shell: $2n^2 = 2(4)^2 = 2\times16 = 32$

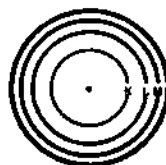



Figure showing different energy levels or shells counted from the centre outwards.

b. Sub-shell:

"Each shell consists of one or more sub shells or orbitals. Each subshell is designated by a small alphabet called letter s, p, d, f etc."

Properties of subshells:

i. First energy level or K shell has only one subshell called s subshell.

ii. Second energy level L, shell has two subshells s and p.

iii. Third energy level M shell has three subshells s, p and d.

iv. Fourth energy level or N shell has four subshells s, p, d and f.

n value	Shell	Subshell
1	K	Only s
2	L	s, p
3	M	s, p ,d
4	N	s, p, d, f

Q.9 What do you mean by electronic configuration? What are basic requirements while writing electronic configuration of an element.

OR

Explain electronic configuration and give the rules for electronic configuration.

Ans: Electronic Configuration:

"The distribution of electrons around the nucleus in various shells and subshells according to their increasing energy is called electronic configuration."

Principle:

The electronic configuration of an atom can be written by using the Aufbau principle.

"In filling the subshells, electrons always enter in lower energy subshell first."

The increasing order in which the electrons will enter into subshell is:

$1s^2$, $2s^2$, sp^6, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$ etc.

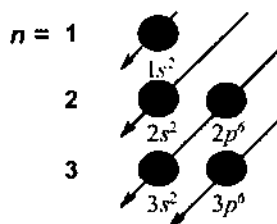
Where,

i. 'n' represents the shell number

ii. Letters (s and p) represent subshells

iii. Subscript shows the number of electrons in a subshell.

The sum of subscripts number is the total number of electrons in an atom i,e. atomic number of an element. as following:

$1s^2$, $2S^2$, $2p^6$, $3s^2$, $2p^6$.............

Rules for electronic configuration

i. The most stable or ground state electronic configuration of an atom is the one in which electrons are present in the lowest possible energy level.

ii. The electrons fill the shells in order of their increasing energy, i.e. lower energy level is occupied first then the higher energy level and so on as indicated earlier.

The maximum capacity of sub shells to accommodate electrons:

's' orbital can accommodate 2 electrons.
'p' orbital can accommodate 6 electrons.
'd' orbital can accommodate 10 electrons.
'f' orbital can accommodate 14 electrons.

The maximum capacity of shells to accommodate electrons:

The maximum capacity of shells to accommodate electrons is as follows:

K shell: $2n^2 = 2(1)^2 = 2\times1 = 2$
L shell: $2n^2 = 2(2)^2 = 2\times4 = 8$
M shell: $2n^2 = 2(3)^2 = 2\times9 = 18$
N shell: $2n^2 = 2(4)^2 = 2\times16 = 32$

As we know there is a slight difference between the energies of the sub shells or orbital within a shell, therefore, filling of electrons in sub shells of a shell is such as that's' sub shell is filled first and then its p sub shell and then other sub shells are filled.

Basic requirements for writing electronic configuration:

While writing the electronic configuration of the elements and their ions, we should know three things.

i. The number of electrons in an atom or ion.
ii. The sequence of shells and subshells according to the energy levels.
iii. The maximum number of electrons that can be placed in different shells and sub shells.

Q.10 Draw electronic configuration of first 18 elements.

Ans: The electronic configuration of first 18 elements

Element	Symbol	Atomic Number	Electronic Configuration
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2, 2s^1$
Beryllium	Be	4	$1s^2, 2s^2$
Boron	B	5	$1s^2, 2s^2, 2p^1$
Carbon	C	6	$1s^2, 2s^2, 2p^2$
Nitrogen	N	7	$1s^2, 2s^2, 2p^3$
Oxygen	O	8	$1s^2, 2s^2, 2p^4$
Fluorine	F	9	$1s^2, 2s^2, 2p^5$
Neon	Ne	10	$1s^2, 2s^2, 2p^6$
Sodium	Na	11	$1s^2, 2s^2, 2p^6, 3s^1$
Magnesium	Mg	12	$1s^2, 2s^2, 2p^6, 3s^2$
Aluminium	Al	13	$1s^2, 2s^2, 2p^6, 3p^1$
Silicon	Si	14	$1s^2, 2s^2, 2p^6, 3p^2$
Phosphorus	P	15	$1s^2, 2s^2, 2p^6, 3p^3$

Sulphur	S	16	$1s^2, 2s^2, 2p^6, 3p^4$
Chlorine	Cl	17	$1s^2, 2s^2, 2p^6, 3p^5$
Argon	Ar	18	$1s^2, 2s^2, 2p^6, 3p^6$

Q.11 Define isotopes. Explain the isotopes of hydrogen.

OR

What is an isotope? Describe the isotopes of hydrogen with diagram.

Ans: Isotopes

"The atoms of an element that have same atomic number but different mass numbers are called isotopes."

Properties of isotopes:

i. They have same electronic configuration and number of protons

ii. They differ in the number of neutrons.

iii. Isotopes have similar chemical properties because they depend upon electronic configuration.

iv. They have different physical properties because these depend upon atomic masses.

v. Most of the elements show isotopes.

vi. All isotopes of an element occupy same position in the periodic table.

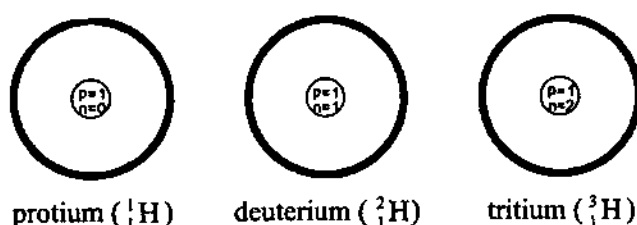
Example:

Isotopes of Hydrogen:

The naturally occurring hydrogen is combination of its three isotopes, present in different abundances. The isotopes of hydrogen are as follows:

The isotopes are represented as:

protium (^1_1H)

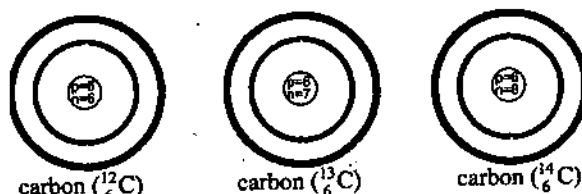
deuterium (^2_1H)

tritium (^3_1H)

	Isotopes	Natural Abundance	At No	Mass No	No of proton	No of Electron	No of Neutron
(i)	Protium (^1_1H or P)	99.985%	1	1	1	1	0
(ii)	Deutrium (^2_1H or D)	0.015%	1	2	1	1	1
(iii)	Tritium (^3_1H or T)	In trace amount	1	3	1	1	2

Q.12 Explain isotopes of carbon, chlorine and uranium.

Ans:

a. Isotopes of Carbon:

There are two stable isotopes of carbon ^{12}C and ^{13}C and one radioactive isotope ^{14}C. Natural abundance of isotopes. The isotope ^{12}C is present in abundance of 98.9%, while ^{13}C and ^{14}C are both present only 1.1 % in nature. All of them have the same number of protons and electrons but differ in number of neutrons. The isotopes are represented as:

b. Isotopes of Chlorine:

There are two isotopes of chlorine $^{35}_{17}Cl$ and $^{37}_{17}Cl$.

Natural abundance:

The isotope Cl-35 is present in abundance of 75% while Cl-37 is present in abundance of 25%. All of them have same number of protons and electrons but differ in number of neutrons.

c. Isotopes of Uranium

There are 3 isotopes of uranium i.e. $^{234}_{92}U$, $^{235}_{92}U$ and $^{238}_{92}U$

The $^{234}_{92}U$ is found in nature nearly 99%.

All of them have same number of protons and electrons but different number of neutrons.

The difference in their number of electrons, protons and neutrons is shown below:

Table: 2.2 Atomic number, mass number, Number of protons and neutrons of H, C, Cl and U

Symbol	Atomic Number	Mass Number	No. of Proton	No. of Neutron
$_1H$	1	1	1	0
$_2H$	1	2	1	1
$_3H$	1	3	1	2
$_{12}C$	6	12	6	6
$_{13}C$	6	13	6	7
$_{14}C$	6	14	6	8
$_{35}Cl$	17	35	17	18
$_{37}Cl$	17	37	17	20
$2_{34}U$	92	234	92	142
$_{235}U$	92	235	92	143
$_{238}U$	92	238	92	146

Q.13 Give the applications of isotopes in the field of radiotherapy, medicines, archaeology, structure determination and power generation.

Ans: Uses or Applications of isotopes:

The major fields in which isotopes have vast applications are the following:

i. Radiotherapy (Treatment of Cancer)

For the **treatment of skin cancer**, isotopes like P-32 and Sr-90 are used because they emit less penetrating beta radiations.

For **cancer, affecting inside the body** Co-60, is used because it emits strongly penetrating gamma rays and beta rays.

ii. Tracer for Diagnosis and Medicine

The radioactive isotopes are used as tracers in medicine to diagnose the presence of tumor in the human body.

Examples:

i. Isotopes of Iodine-131 are used for diagnosis of goiter in thyroid gland.

ii. Similarly technetium is used to monitor the bone growth.

iii. Archaeological and Geological Uses

The radioactive isotopes are used to estimate the age of fossils like dead plants and animals and stones etc.

"The age determination of very old objects based on the half-lives of the radioactive isotope is called radioactive-isotope dating."

"Age determination of old carbon containing objects (fossils) by measuring the radioactivity of C-14 in them is called radio-carbon dating or simply carbon dating. This is an important method of age determination of old objects"

iv. Chemical Reaction and Structure Determination

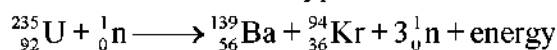
The radioisotopes are used in a chemical reaction to follow a radioactive element during the reaction and ultimately to determine the structure.

C-14 is used to label CO_2. As CO_2 is used by the plants for photosynthesis to form glucose, its movement is detected through the various intermediate steps up to glucose.

v. Applications in Power Generation

Radioactive isotopes are used to generate electricity by carrying out controlled nuclear fission reactions in nuclear reactors.

When U-235 is bombarded with slow moving neutrons, the uranium nucleus breaks up to produce Barium-139 and Krypton and three neutrons

$$^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{139}_{56}Ba + ^{94}_{36}Kr + 3^{1}_{0}n + \text{energy}$$

A large amount of energy is released which is used to convert water into steam in boilers. The steam then drives the turbines to generate electricity. This is the peaceful use of atomic energy for development of a nation.

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 2: Structure of Atoms
Extra MCQ's

1. **Matter is composed of tiny indivisible particles called:**
   (a) Element (b) Atom (c) Compound (d) substance
2. **Atom of the same elements are**
   (a) Different (b) Alike (c) Comparable (d) Active
3. **Gas discharge tube experiment was performed by:**
   (a) J.J. Thomson (b) Rutherford (c) Dalton (d) William Crooks
4. **The pressure inside the discharge tube for the discovery of electron was kept**
   (a) 10^{4} atm (b) 10^{-4} atm (c) 10^{14} atm (d) 10^{-14} atm
5. **Who was the pioneer of the vacuum tubes?**
   (a) William Crooks (b) Rutherford (c) Bohr (d) Dalton
6. **The nature of canal rays depends upon:**
   (a) Nature of anode (b) Nature of Cathode
   (c) Nature of gas (d) Nature of particles
7. **The mass of proton is ______ times more than that of an electron:**
   (a) 1830 (b) 1840 (c) 2 (d) 3
8. **Which one of the followings is produced by the bombardment of the helium particle on beryllium?**
   (a) Alpha particle (b) Beta particle (c) Neutron (d) Gamma rays
9. **The highly penetrating rays are:**
   (a) Alpha particle (b) Beta particle (c) Neutron (d) Both a & b
10. **Neutron was discovered by**
    (a) Rutherford (b) Chadwick (c) Bohr (d) William Crooks
11. **In α -scattering experiment Rutherford used the foil made up of:**
    (a) Silver (b) Tin (c) Platinum (d) Gold
12. **Alpha particles are emitted by radioactive element:**
    (a) Carbon (b) Polonium (c) Neon (d) Vanadium
13. **Rutherford used the photographic plate coated with**
    (a) Zinc sulphide (b) Zinc sulphite (c) Zinc oxide (d) Zinc sulphate
14. **Who is the father of nuclear chemistry?**
    (a) Rutherford (b) Dalton (c) William Crooks (d) Joseph Proust
15. **According to quantum theory which type of spectrum is shown?**
    (a) Continuous spectrum (b) Line spectrum
    (c) Emission spectrum (d) Absorption spectrum
16. **Which of the following are fundamental particles of an atom?**
    (a) Ion (b) Molecular ion (c) Electron (d) Positron
17. **Matter is composed of tiny indivisible particles called:**
    (a) Ion (b) Free radical (c) Atoms (d) Molecules
18. **The meaning of Latin word 'atom' is:**
    (a) Chroma (b) Divisible (c) Atomos (d) Same place
19. **Rutherford used a thin sheet of gold of thickness:**
    (a) 0.00004cm (b) 0.004cm (c) 0.0004cm (d) 0.04cm
20. **Canal rays were discovered by:**

(a) Goldstein (b) Thomson (c) Dalton (d) William Crooks

21. **Protons were discovered by:**
(a) Thomson (b) Chadwick (c) Moseley (d) Goldstein

22. **Rutherford bombarded a thin sheet of gold with:**
(a) α-particles (b) β-particles (c) γ-particles (d) x-rays

23. **Which apparatus was used by Sir William Crooks in his experiment?**
(a) Test tube (b) Gas discharge tube (c) Zinc plate (d) Electrolytic cell

24. **Which are three fundamental particles of an atom?**
(a) Ion, radicals, free radicals (b) Electrons, protons, neutrons
(c) Electrons, protons, cathode rays (d) Canal rays, x-rays, gamma rays

25. **The electrons revolve around the:**
(a) Atom (b) Nucleus (c) Protons (d) Neutrons

26. **Nature of gas present in discharge tube affects the nature of:**
(a) Canal rays (b) x-rays (c) Cathode rays (d) β-rays

27. **Alpha particles are basically nucleus of:**
(a) Lithium (b) Sodium (c) Potassium (d) Helium

28. **Plum pudding model was put forwarded by:**
(a) Dalton (b) Thomson (c) Goldstein (d) Chadwick

29. **Neil Bohr won the noble prize in:**
(a) 1914 (b) 1918 (c) 1922 (d) 1926

30. **Canal rays travel in a straight line in a direction _______ to cathode rays.**
(a) Opposite (b) Same (c) Parallel (d) None of these

31. **Rutherford, Moseley, Bohr and other scientists performed experiments and revealed that:**
(a) Atom has complex nature (b) Atom is neutral
(c) Atom can be divisible (d) Atom is beyond understanding

32. **Canal rays carry:**
(a) +ve charge (b) –ve charge (c) Neutral (d) None of these

33. **Which isotope of carbon is in abundance?**
(a) ^{12}C (b) ^{13}C (c) ^{14}C (d) Both a and b

34. **Isotopes have different number of :**
(a) Electron (b) Proton (c) Neutron (d) Charge

35. **Energy of N-shell is more than the energies of:**
(a) K (b) K, L (c) K, L, M (d) L, M

36. **Quantum means:**
(a) Fixed volume (b) Fixed energy (c) Fixed pressure (d) Fixed temperature

37. **The subshells of M-shell are:**
(a) s, p (b) s, p, d (c) s, p, d, f (d) s, d, f

38. **N-shell can accommodate a maximum of ____ electrons.**
(a) 8 (b) 2 (c) 18 (d) 32

39. **Which one is the electronic configuration of Cl^{-1}?**
(a) $1s^2, 2s^2, 2p^6, 3s^1$ (b) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$
(c) Both a and b (d) None of these

40. **Which one is the electronic configuration of sulphur?**
(a) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$, (b) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$
(c) $1s^2, 2s^2, 2p^6, 3s^2, 3p^4$ (d) $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

41. **The value of Planck's constant is:**
(a) 6.62×10^{-34} Js (b) 6.62×10^{-24} Js (c) 6.62×10^{-19} Js (d) 6.62×10^{-12} Js

42. **Which one of the followings results in the discovery of proton?**

(a) Cathode rays (b) Canal rays (c) x-rays (d) Alpha rays

43. The concept of orbit was introduced by:
(a) J.J. Thomson (b) Rutherford (c) Bohr (d) Planck

44. Deutrium is used to make:
(a) Light water (b) Heavy water (c) Soft water (d) Hard water

45. Co-60 is the source of:
(a) X-rays (b) Beta radiations (c) Alpha particles (d) Gamma rays

46. For the production of cathode rays the pressure of gas inside the discharge tube was Kept:
(a) 10^{-1}atm (b) 10^{-2}atm (c) 10^{-4}atm (d) 10^{-5}atm

47. Which one of the shells contains f–subshells?
(a) K (b) L (c) M (d) N

48. Which one of the followings has only one neutron in its nucleus?
(a) Protium (b) Deutrium (c) Tritium (d) Helium

49. Beta radiations are emitted by:
(a) Co-60 (b) C-12 (c) S-16 (d) Sr-90

50. Rutherford won nobel prize in:
(a) 1909 (b) 1906 (c) 1908 (d) 1910

51. ${}^{9}_{4}Be + {}^{4}_{2}He \longrightarrow {}^{12}_{6}C + ?$
(a) ${}^{1}_{0}p$ (b) ${}^{-1}_{0}e$ (c) ${}^{1}_{0}n$ (d) ${}^{4}_{2}He$

ANSWER KEY

1	b	12	b	23	b	34	c	45	d
2	b	13	a	24	b	35	c	46	c
3	d	14	a	25	b	36	b	47	d
4	b	15	b	26	a	37	b	48	b
5	a	16	c	27	d	38	d	49	d
6	c	17	c	28	b	39	b	50	c
7	b	18	a	29	c	40	c	51	c
8	c	19	a	30	a	41	a		
9	c	20	a	31	c	42	b		
10	b	21	d	32	a	43	c		
11	d	22	a	33	a	44	b		

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 3 – Periodic Table and Periodicity of Properties
Exercise Questions

Exercise Multiple Choice Question Answers:

1. **The atomic radii of the elements in Periodic Table:**
(a) Increase from left to right in a period (b) Increase from top to bottom in a group
(c) Do not change from left to right in a period (d) Decrease from top to bottom in a group
2. **The amount of energy given out when an electron is added to an atom is called:**
(a) Lattice energy (b) ionization energy (c) Electronegativity (d) Electron affinity
3. **Mendeleev Periodic Table was based upon the:**
(a) Electronic configuration (b) atomic mass
(c) Atomic number (d) completion of a subshell
4. **Long form of Periodic Table is constructed on the basis of:**
(a) Mendeleev Postulate (b) Atomic number
(c) Atomic mass (d) Mass number
5. **4th and 5^{th} period of the long form of Periodic Table are called:**
(a) Short periods (b) normal periods (c) Long periods (d) Very long periods
6. **Which one of the following halogen has lowest electronegativity?**
(a) Flourine (b) chlorine (c) Bromine (d) Iodine
7. **Along the period, which one of the following decreases:**
(a) Atomic radius (b) ionization energy (c) Electron affinity (d) Electronegativity
8. **Transition elements are:**
(a) All gases (b) all metals (c) All non-metals (d) All metalloids
9. **Mark the incorrect statement about ionization energy:**
(a) It is measured in $kJmol^{-1}$ (b) It is absorption of energy
(c) It decreases in a period (d) It decreases ina group
10. **Point out the incorrect statement about electron affinity:**
(a) It is measured in $klmol^{-1}$ (b) It involves release of energy
(c) It decreases in a period (d) It decreases in a group

ANSWR KEY

1	b	3	b	5	c	7	a	9	c
2	d	4	b	6	d	8	b	10	c

Exercise Short Questions Answers

Q.1 Why noble gases are not reactive?

Ans: Noble gases are not reactive because they have their valence shells completely filled. They have 2 or 8 electrons in their valence shells. Their atoms do not have vacant spaces in their valence shell to accommodate more electrons. Therefore they do not gain, lose or share electrons.

Q.2 Why Cesium (at.no.55) requires little energy to release its one electron present in the outermost shell?

Cesium requires little energy because it has greater atomic size, more shielding effect (due to presence of more electrons) and low ionization energy due to which the hold of inner nucleus on valence.

Q.3 How is periodicity of properties dependent upon number of protons in an atom?

Ans: Number of protons in an atom represents atomic number of that element which increases regularly by one form element to element. So the arrangement of elements according to increasing atomic number shows the periodically in the electronic configuration of the elements that leads to periodicity in their properties.

Q.4 Why shielding effect of electrons makes cation formation easy?

Ans: The shielding effect of electrons makes the cation formation easy because it reduces the nuclear pull on the outermost electrons and they are less tightly held by the nucleus and can easily be lost from the outermost shell.

Q.5 What is the difference between Mendeleev's periodic law and modem periodic law?

Ans:

Mendeleevs periodic law	Modern periodic law
Properties of the elements are periodic function of their atomic masses.	Properties of the elements are periodic function of their atomic numbers.
Atomic masses is less fundamental property and it is the basis of mendeleevs periodic law.	Atomic number is more fundamental property and it is the basis of modern periodic law.

Q.6 What do you mean by groups and periods in a Periodic Table?

Ans: The horizontal rows of elements in a periodic table are called periods. The vertical columns in a periods table are called group. There are 18 groups in the long form of the periodic table. They are studied from top to bottom.

Q.7 Why and how are elements arranged in 4th period?

Ans: The elements (Na, Mg, Al, Si, P, S, Cl and Ar) are arranged in the 4th period because they are all having four electronic shells and are arranged by increasing atomic number

from left to right the period.

Q.8 Why the size of atom does not decrease regularly in a period?

Ans: The size of atom does not decrease regularly in a period. This irregularity in the transition metals is due to the involvement of d orbital. It provides poor shielding effect.

Q.9 Give the trend of ionization energy in a period.

Ans: ionization energy increases from left to right in a period and decreases from top to bottom in a group.

Reason:

It is because the size of atoms reduces and valence electrons are held strongly by the electrostatic force of nucleus.

Exercise Long Question Answers

Q.1 Explain the contributions of Mendeleev for the arrangement of elements in a Periodic Table.

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Q.2 Show why in a 'period' the size of an atom decreases if one moves from left to right.

Ans: See Q. No. **11** (Subjective Part, Long Questions Answers)

Q.3 Describe the trends of electronegativity in a period and in a group.

Ans: See Q. No. **15** (Subjective Part, Long Questions Answers)

Q.4 Discuss the important features of modem Periodic Table.

Ans: See Q. No. 7 (Subjective Part, Long Questions Answers)

Q.5 What do you mean by blocks in a periodic table and why elements were placed in blocks?

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.6 Discuss in detail the periods in Periodic Table?

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.7 Why and how elements are arranged in a Periodic Table?

Ans: See Q. No. **5** (Subjective Part, Long Questions Answers)

Q.8 What is ionization energy? Describe its trend in the Periodic Table?

Ans: See Q. No. **13** (Subjective Part, Long Questions Answers)

Q.9 Define electron affinity, why it increases in a period and decreases in a group in the Periodic Table.

Ans: See Q. No. **14** (Subjective Part, Long Questions Answers)

Q.10 Justify the statement, bigger size atoms have low ionization energy and have more shielding effect.

Ans: Ionization Energy:

The amount of energy required to remove the most loosely bound electron from the valence shell of an isolated gaseous atom is called ionization energy."

Shielding Effect:

"The decrease in the attractive force exerted by the nucleus on the valence shell electrons due to the presence of the electrons lying between the nucleus and valence shell is called shielding effect."

As we move down the group more and more shells lie between the valence shell and the nucleus of the atom, these additional shells reduce the electrostatic force felt by the electron present in the outermost shell which results more shielding effect by such bigger size atoms. Resultantly the valence shell electrons can be released easily. Therefore bigger size atoms have more shielding effect and low ionization energies.

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 3: Periodic Table and Periodicity of Properties
Long Questions

Q.1 Describe Dobereiner's triads with the help of an example.

Ans: Dobereiner's Triads

Introduction:

A German chemist Dobereiner observed relationship between atomic masses of several groups of three elements called triads.

Law of triads:

"In a triad the central or middle element had atomic mass average of the other two elements."

Example:

One triad group example is that of calcium (40), strontium (88) and barium (137). The atomic mass of strontium is the average of the atomic masses of calcium and barium.

Drawbacks:

i. Only a few elements could be arranged in this way.

ii. This classification did not get wide acceptance.

Cannizzaro:

He successfully determined the correct atomic masses of elements in 1860.

Q.2 Write a note on Newlands octaves.

Ans: Newlands Octaves

Introduction:

In 1864 British chemist and musician Newlands put forward his observations in the form of 'law of octaves'.

Statement

"According to law of Octaves there was a repetition in chemical properties of every eighth element if they were arranged by their increasing atomic masses."

He compared it with musical notes.

Drawbacks

i. His work could not get much recognition as no space was considered for undiscovered elements.

ii. The noble gases were also not known at that time.

Q.3 Explain the contributions of Mendeleev for the arrangement of elements in his periodic table.

Ans: Mendeleev's Periodic Table

Introduction:

A Russian chemist, Mendeleev arranged the known elements (only 63) in order of increasing atomic masses, in horizontal rows called periods, so that elements with similar properties were in the same vertical columns. This arrangement of elements was called Periodic Table.

Mendeleev's Periodic law,

"Properties of the elements are periodic functions of their atomic masses"

Demerits of Mendeleev's periodic table:

i. It did not explain the position of isotopes.
ii. Wrong order of the atomic masses of some elements suggested that atomic mass of an element cannot serve as the basis for the arrangement of elements.

Q.4 What is modern periodic law and modern periodic table?

Ans: Modern Periodic Law

Introduction:

In 1913 H. Moseley discovered a new property of the elements i.e. atomic number. He observed that atomic number instead of atomic mass should determine the position of elements in the periodic table

Periodic Law:

"Properties of the elements are periodic function of their atomic numbers".

Note:

Atomic number of an element is equal to the number of electrons in neutral atom. So atomic number provides the basis of electronic configuration as well.

Q.5 Modern periodic table is based upon atomic number and periodicity?

Ans: Modern Periodic Table:

"A table obtained by arrangement of elements into groups and periods in increasing order of their atomic number is called modern periodic table."

Atomic number of an element is more fundamental property than atomic mass

i. It increases regularly from element to element.
ii. It is fixed for every element.

So the discovery of atomic number of an element in 1913 led to change in Mendeleev's periodic law which was based on atomic mass.

Basis of Modern Periodic Table:

The modern periodic table is based upon the arrangement of elements according to increasing atomic number. When the elements are arranged according to increasing atomic number from left to right in a horizontal row, properties of elements were found repeating after regular intervals such that elements of similar properties and similar configuration are placed in the same group. It was observed that after every eighth element, ninth element had similar properties to the first element.

Example:

Sodium (Z=11) had similar properties to lithium (Z=3). After atomic number 18, every nineteenth element was showing similar behaviour. So the long rows of elements were cut into rows of eight and eighteen elements, and placed one above the other so that a table of vertical and horizontal rows was obtained.

Q.6 Why and how elements are arranged in the periodic table?

OR

What is the significance of atomic number in modern periodic table?

Ans: Long form of periodic table:

Long form of the periodic table was first proposed by a Danish chemist Tulius Thomson in 1895. It is also known as Bohr's periodic table. It was developed by Rang modified by Warner and extended by Bury. The periodic table is arranged in 7 horizontal rows called periods and 18 vertical columns called groups.

The significance of atomic number

The significance of atomic number in the arrangement of elements in the modem periodic table lies in the fact that as electronic configuration is based upon atomic number, so the arrangement of elements according to increasing atomic number shows the periodicity (repetition of properties after regular intervals) in the electronic configuration of the elements that leads to periodicity in their properties. Hence the arrangement of elements based on their electronic configuration created a long form of periodic table

Periods:

Definition

"The horizontal rows of elements in a periodic table are called periods."

The first element of each period is an alkali metal while the last element is noble gas.

Properties:

i. The elements in a period have continuously increasing atomic number i.e. continuously changing electronic configuration along a period.

ii. As a result properties of elements in a period are continuously changing.

iii. The number of valence electrons decides the position of an element in a period.

Examples:

i. Elements which have 1 electron in their valence shell occupies the left most position in the respective periods, such as alkali metals.

ii. Similarly the elements having 8 electrons in their valence shells such as noble gases always occupy the right most position in the respective periods.

Groups:

The vertical columns of elements in the periodic table are called groups.

Properties:

i. These groups are numbered from left to right as 1 to 18.
ii. The elements in a group do not have continuously increasing atomic numbers.
iii. Rather the atomic numbers of elements in a group increase with irregular gaps.
iv. But the elements of a group have similar electronic configuration i.e. same number of electrons are present in the valence shell.

Examples:

i. The first group elements have only 1 electron in their valence shells.
ii. Similarly group 2 elements have 2 electrons in their valence shells.

Note:

It is the reason elements of a group have similar properties.

Q.7 Discuss the important features of Long Form of Periodic Table.

Salient Features of Long Form of Periodic Table

The salient features of long form of periodic table are as follows:

Periods:

i. This table consists of seven horizontal rows called periods
ii. First period consists of only two elements. Second and third period consist of 8 elements each. Fourth and fifth period consist of 18 elements each. Sixth period has 32 elements while seventh period has 32 elements and is incomplete.
iii. Elements of a period show different properties.

Groups:

i. There are 18 vertical columns in the periodic table numbered 1 to 18 from left to right, which are called groups.
ii. The elements of a group show similar properties.

Blocks:

i. Elements are classified into four blocks depending upon the type of the sub- shell which gets the last electron.

Fig. 3.1 Modern Periodic Table or long form of the Periodic Table of Elements.

Q.8 What do you mean by blocks in the periodic table and why elements were placed in blocks?

Blocks of elements:

On the basis of completion of a particular sub shell, elements with similar sub shell electronic configuration are referred as a block of elements.

Types of blocks:

- **i.** There are four blocks in the periodic table named after the name of the sub shell which is in the process of completion by the electrons.
- **ii.** These are s, p, d and f blocks in the periodic table.

a. s-block:

The elements in which valence electrons are present in the s-subshell are called s-block elements."

"Elements of group 1 and 2 have valence electrons in 's' subshell. Therefore, they are called s-block elements."

b. p-block:

The elements in which valence electrons are present in the p-subshell are called p-block elements."

"Elements of group 13 to 18 have their valence electrons in 'p' sub shell. Therefore, they are referred as p-block elements."

c. d-block:

The elements in which valence electrons are present in the d-subshell are called d-block

elements."

"The elements of group 3 to group 12 have their valence electrons in d subshell. Therefore they are called d-block elements.

The d-block constitutes period 4, 5 and 6. Each period in d-block consists of ten groups starting from group 3 to group 12. These are called transition metals."

d. f-block:

The elements in which valence electrons are present in the f-subshell are called f-block elements."

"f-block lies separately at the bottom of the periodic table. It consists of lanthanides and actinides.

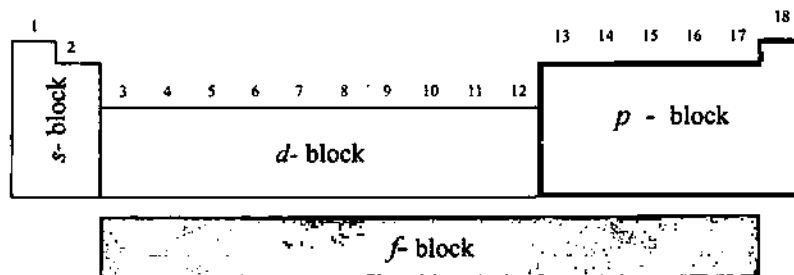



Fig. 3.2 Modern Periodic Table showing four blocks.

DO YOU KNOW

Alchemists:

Group of Muslim scientists who tried to convert cheaper metals into gold and to give eternal life to the people is called alchemist and this branch of chemistry is called alchemy.

Alchemy:

For thousand of years alchemy remained field of interest for the scientists.

Main objective of Alchemists:

They worked with two main objectives; change common metals into gold and second find cure to diseases and give eternal life to people. They believed all kinds of matter were same combination of four basic elements. Substances are different because these elements combine differently. Changing composition or ratio of anyone element, new substances can be formed.

Disadvantages and Advantages:

The way of making gold from silver or lead was never found and secret of eternal life was never discovered. However, many methods and techniques invented by alchemists are still used in chemistry.

Q.9 Write in detail the periods of periodic table.

Ans: Periods:

"Horizontal rows of elements in the periodic table are called periods."

There are seven periods in the modern periodic table. The period number of an element represents number of shells in the element.

First period:

It is called short period. It consists of only two elements, hydrogen and helium.

Second and third periods:

These are called normal periods. Each of them has eight elements in it. Second period consists of lithium beryllium, boron, carbon, nitrogen, oxygen, fluorine and ends at neon, a noble gas.

Fourth and fifth periods:

These are called long periods. Each one of them consists of eighteen elements.

Sixth and seventh periods:

These are called very long periods. Sixth period contains 32 elements whereas seventh period is incomplete.

Lanthnides and actinides:

In sixth and seventh period after atomic number 57 and 89, two series of fourteen elements each, were accommodated.

a. Why lanthanides and actinides are placed separately?

Because of space problem, these two series were placed separately below the normal periodic table to keep it in a manageable and presentable form.

b. Why lanthanides and actinides are called so?

Since the two series start after Lanthanum (Z=57) and Actinium (Z=89), so these two series of elements are named as Lanthanides and Actinides respectively.

Starting and ending of a period:

All the periods, except the first period start with an alkali metal and end at a noble gas. It is to be observed that number of elements in a period is fixed because of maximum number of electrons which can be accommodated in the particular valence shell of the elements.

Period No.	Name of the Period	Number of Elements	Range of Atomic Numbers
1st	Short Period	2	1 to 2
2nd	Normal Period	8	3 to 10
3rd		8	11 to 18
4th	Long Period	18	19 to 36
5th		18	37 to 54
6th	Very Long Period	32	55 to 86
7th		32*	87 to 118*

Q.10 Write a detailed note on the groups of periodic table.

Ans:

Groups:

"The vertical columns of elements in the periodic table are called groups."

Group 1: Consists of hydrogen, lithium, sodium, potassium, rubidium, cesium and francium.

Although elements of a group do not have continuously increasing atomic numbers, yet

they have similar electronic configuration in their valence shells.

Family Name:

Elements of group are also called family because normal elements of a group have similar chemical properties and similar electronic configuration in their valence shells.

Important Groups:

Group 1: It consists of hydrogen (H), Lithium (Li), Sodium (Na), potassium (K), rubidium (Rb), cesium (Cs) and francium. They are generally called alkali metals.

Group 2: It consists of beryllium (Be), magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra). They are called alkaline earth metals.

Group 17: It consists of Fluorine (F), Chlorine (Cl), Bromine (Br), Iodine (I) and astatine (At). The elements of this group are called halogens.

Group 18: The gaseous elements of group 18 or zero group are called noble gases. It consists of helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe) and Radon (Rn). The elements of this group are called noble gases.

Groups of Normal elements: (Representative or typical elements)

All s-block and p-block elements excluding noble gases are called normal elements. The groups 1,2 and 13 to 17 contain the normal elements. In the normal elements all the inner shells are completely filled with electrons, only the outermost shells are incomplete.

For example, group 17 elements (halogens) have 7 electrons in their outermost (valence) shell.

Transition elements:

"Elements in which 'd' or 'f' subshell are in the process of completion are called transition elements."

The elements of groups 3 to 12 and lanthanides as well as actinides are called transition elements. They belong to periods 4,5, 6 and 7.

Valence electrons	Group number	Family name	General Electronic configuration
1 electron	1	Alkali metals	ns^1
2 electrons	2	Alkaline earth metals :	ns^2
3 electrons	13	Boron family	$ns^2\ np^1$
4 electrons	14	Carbon family	$ns^2\ np^2$
5 electrons	15	Nitrogen family	$ns^2\ np^3$
6 electrons	16	Oxygen family	$ns^2\ np^4$
7 electrons	17	Halogen family	$ns^2\ np^5$
8 electrons	18	Noble gases	$ns^2\ np^6$

Do You Know

Fireworks

Beautiful fireworks display are common on celebrations like Pakistan Day or even on marriages. It is dangerous but careful use of various elements and particularly metal salts of different composition give beauty and colors to the fireworks.

Invention of fireworks:

A technology invented in China is used all over the world.

Composition of fireworks:

Elements like magnesium, aluminium are used in powdered form. Usually nitrates and chlorates are used. Other chemicals are added to give brilliance and different shades.

Salt	Colour Imparted
Sodium salt	Yellow
Calcium salt	Red
Strontium salt	Scarlet
Barium salt	Green
Copper salt	Bluish green

Precautions:

Because of fire hazard and risk to life and property, only skilled professionals use them.

Q.11 What is meant by atomic size? Give its units of measurements and explain its trends in modern periodic table.

Ans: Atomic size or atomic radius:

"*The half of the distance between the nuclei of the two bonded atoms is referred as the atomic radius of the atom.*

Example:

The distance between the nuclei of two carbon atoms in its elemental form is 154 pm, it means its half 77 pm is radius of carbon atom.

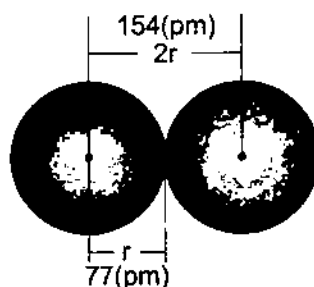



Fig. 3.3 The radius of carbon atom.

Units:

The units of atomic radius are as follows.

i. Nanometer (10^{-9}m)

ii. Picometer (10^{-12}m)

iii. Angstrom (10^{-10}m)

Trends of atomic radius and atomic size in periods:

The atomic radii gradually decrease from left to right in a period.

Reasons:

It is because with the increase of atomic number, the effective nuclear charge increases gradually because of addition of more and more protons in the nucleus. This nuclear force pulls down or contracts the outermost shell towards the nucleus.

Example:

Atomic size in period 2 decreases from Li (152 pm) to Ne (69 pm)

Trend of Atomic Radius and Atomic size in Groups:

The atomic radii increase from top to bottom in a group.

Reason:

The number of shells increases in the successive elements. The distance between the nucleus and valence shells increases, the effective nuclear charge decreases and atomic radius increases.

Q.12 What is shielding effect? Write down its trend in modern periodic table.

Ans: Shielding Effect

"The decrease in attractive force exerted by the nucleus on the valence shell electrons due to the presence of electrons lying between the nucleus and valence shell is called shielding effect."

Effective nuclear charge

The attraction of outer electrons toward nucleus is partially reduced because of presence of inner electrons. As a result an atom experiences less nuclear charge than that of the actual charge, which is called effective nuclear charge (Z_{eff}).

Explanation:

The shielding effect decreases the forces of electrostatic attractions between nucleus and outermost electrons by partially concealing or blocking the nuclear attraction for the outer most electrons. In fact the electrons present between the nucleus and the outer most shell of an atom reduce the effective nuclear charge felt by the electrons present in the outermost shell.

Trend of shielding effect in groups:

The shielding effect increases down the group in the periodic table.

Reason:

This is because the number of inner shells increases from top to bottom in a group.

For example:

Due to greater size of the atom it is easy to take away electron from potassium (z=19) than from sodium (z =11) atom.

Trends of shielding effect in periods:

The shielding effect does not change in a period.

Reason:

This is because the number of inner shells remain the same from left to right in the periods.

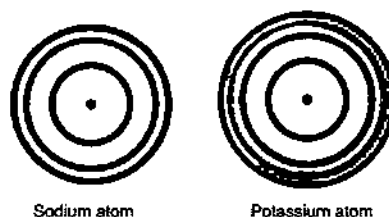



Fig. 3.4: Shielding effect is more in potassium atom than that of sodium atom.

Q.13 What is ionization energy? Describe its trends in modern periodic table.

Ans: Ionization Energy (I.E)

"The amount of energy required to remove the most loosely bound electron from the valence shell of an isolated gaseous atom is called ionization energy."

Units of I.E:

The units of ionization energy are KJ mol^{-1} and ev/atom.

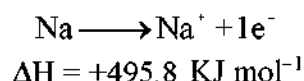
First Ionization Energy:

The amount of energy required to remove the first electron from the valence shell of an isolated gaseous atom is called first ionization energy.

Example:

The first ionization energy of sodium atom is + 495.8 $kJmol^{-1}$

$$Na \longrightarrow Na^{+} + 1e^{-}$$

$$\Delta H = +495.8 \text{ KJ mol}^{-1}$$

Second ionization energy:

"The amount of energy required to remove the second electron from the valence shell of an isolated gaseous monopositive ion is called second ionization energy."

When there are more than one electrons in valence shell they can be removed one by one providing more and more energy. Such as group 2 and 3 elements have more than one electron in their valence shells. Therefore, they will have more than one ionization energy values.

For example:

$$Mg \longrightarrow Mg^{+} + 1e^{-} \qquad \Delta H = +738 \text{ KJ mol}^{-1}$$

$$Mg^{+1} \longrightarrow Mg^{+2} + 1e^{-} \qquad \Delta H = +1450 \text{ KJ mol}^{-1}$$

Third ionization energy:

"The amount of energy required to remove the third electron from the valence shell of an isolated gaseous di-positive ion is called third ionization energy."

Trends along groups

Ionization energy of elements decreases from top to bottom in a group.

Reason:

i. The number of shells increases

ii. The distance between the nucleus and valence shells increases.

iii. Shelling effect increases.

iv. Nuclear attraction on valence electrons decreases.

Therefore, ionization energy decreases from top to bottom in the groups of the periodic table.

Trend along periods

Ionization energy values of elements increase from left to right in a period

Reason:

i. The number of shells remains same.

ii. Shieding effect remains same.

iii. The effective nuclear charge on valence electrons increases.

iv. The distance between nucleus and valence shells decreases.

v. Nuclear attraction on valence electrons increases.

1st group elements	Ionization (KJ/mol)
$_3Li$	520
$_{11}Na$	496
$_{19}K$	419
$_{37}Rb$	403
$_{55}Cs$	377

Therefore, ionization energy increases from left to right in periods of the periodic table.

2nd period elements	$_3Li$	$_4Be$	$_5B$	$_6C$	$_7N$	$_8O$	$_9F$	$_{10}Ne$
Ionization energy ($kJmol^{-1}$)	520	899	801	1086	1402	1314	1681	2081

Q.14 Define electron affinity? Why it increases in a period and decreases in a group?

Ans: Electron Affinity

"*The amount of energy released when an electron is added up in the outermost shell of an isolated gaseous atom is called electron affinity.*"

Example:

The electron affinity of fluorine is -328 kJ mol^{-1} i.e. one mole atom of fluorine releases 328 kJ of energy to form one mole of fluoride ions.

$$F + 1e^- \longrightarrow F^- \qquad \Delta H = -328 \text{ KJ mol}^{-1}$$

Affinity means attraction. Therefore, electron affinity means tendency of an atom to accept an electron to form an anion.

Units of electron affinity:

The units of electron affinity are $KJmol^{-1}$ and eV/atom.

Trend of electron affinity along period:

Electron affinity values increase from left to right in the period. .

Reason:

The reason for this increase is, as the size of atoms decreases in a period, the attraction of the nucleus for the incoming electron increases. That means more is attraction for the electron, more energy will be released.

Trend of electron affinity along group:

In a group electron affinity values decrease from top to bottom because the size of elements of atoms increases down the group.

Reason:

With the increase in size of atom shielding effect increases that results in poor attraction for the incoming electron i.e. less energy is released out. For example, as the size of iodine atom is bigger than chlorine, its electron affinity is less than chlorine.

2nd period elements	$_3Li$	$_4Be$	$_5B$	$_6C$	$_7N$	$_8O$	$_9F$	$_{10}Ne$
Electron affinity ($kJmol^{-1}$)	-60	>0	-29	-122	0	-141	-328	0

Q.15 What is electronegativity? Write down its trends in modern periodic table.

Ans: Electronegativity:

"The ability of an atom to attract the shared pair of electrons towards itself in a molecule is called electronegativity."

Explanation:

It is an important property especially when covalent type of bonding of elements is under consideration.

Trends in periods:

Electronegativity increases from left to right in the periodic table. The trend of electro negativity is same as of ionization energy and electron affinity. It increases in a period from left to right.

Reason:

Because higher (Z_{eff}) shortens distance from the nucleus of the shared pair of electrons. Thus, enhances the power to attract the shared pair of electrons.

Example:

Electronegativity values of group 17 are given as follows:

Trend in groups:

Electronegativity decreases from top to bottom in the group. It generally decreases down a group because size of the atom increases. Thus attraction for the shared pair of electrons weakens. For example, electronegativity values of group 17 (halogens) are presented here.

17^{th} group elements	Electro negativity
$_9F$	4.0
$_{17}Cl$	3.2
$_{35}Br$	3.0
$_{53}I$	2.7

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 3: Periodic Table and Periodicity of Properties
Extra MCQ's

1. **Shielding effect is greater in atoms with greater number of:**
   (a) Protons (b) Neutrons (c) Electrons (d) Positrons
2. **There are total ____ groups and _____ periods in the modern periodic table:**
   (a) Seven, eight (b) Eighteen, seven (c) Eight, seventeen (d) Sixteen, eight
3. **If we move from left to right in a period, the value of ionization energy:**
   (a) Remains same (b) Decreases (c) Increases (d) Not effected
4. **The ionization energy of sodium is:**
   (a) +495.8 kj/mol (b) +594.8 kj/mol (c) -495.8 kj/mol (d) -594.8 kj/mol
5. **It is the amount of energy released when an electron is added up in the outermost shell of an isolated gaseous atom:**
   (a) Shielding effect (b) Electron affinity (c) Electro negativity (d) Ionization energy
6. **Salts of sodium give:**
   (a) Bluish green (b) Red (c) Yellow (d) Green
7. **The half of distance between the nuclei of the two bonded atoms is referred as:**
   (a) Atomic size (b) Atomic radius (c) Ionic radii (d) Both (a) and (b)
8. **Sixth and seventh periods are called:**
   (a) Short period (b) Normal period (c) Long period (d) Very long period
9. **The d-block elements lie between the blocks:**
   (a) s-p (b) d-f (c) p-s (d) f-d
10. **Which one of the halogens has the highest electro-negativity?**
    (a) Bromine (b) Iodine (c) Chlorine (d) Fluorine
11. **The shielding effect of the inner electrons is responsible for:**
    (a) Increasing ionization energy value (b) Decreasing ionization energy value
    (c) Increasing electron affinity (d) Increasing electronegativity
12. **According to the modern periodic law, the properties of the elements are periodic function of their:**
    (a) Atomic number (b) Number of electrons
    (c) Mass number (d) Number of valence electrons
13. **Which is the best reason for increasing ionization energy from left to right in a period?**
    (a) The shielding effect remains the same (b) The nuclear chare increases
    (c) The number of inner electrons increases (d) Increasing electronegativity
14. **Units of ionization energy are:**
    (a) Calories/mol (b) kj/mol (c) kj (d) j/mol
15. **5-f series of inner transition elements are called:**
    (a) Lanthanides (b) Actinides (c) Halogens (d) Alkali metals
16. **Halogens belong to group:**

(a) 17 (b) 16 (c) 18 (d) 32

17. How many elements are there in 4^{th} period of the periodic table?
(a) 7 (b) 8 (c) 18 (d) 32

18. Which one is the incomplete period in the periodic table?
(a) fourth period (b) second period (c) seventh period (d) first period

19. Which scientist had given the idea of octaves for the arrangement of elements?
(a) Mendeleev (b) Al-Razi (c) Newlands (d) Dobereiner

20. Atomic number of K = 19. Its valence shell configuration is:
(a) $4s^1$ (b) $3s^1$ (c) $6s^1$ (d) $5s^1$

21. Keeping in view the size of atom, which order is the correct one?
(a) Mg > Sr (b) Ba> Mg (c) Ca > Ba (d) Cl > I

22. In 5^{th} period the elements have range from:
(a) 19 to 36 (b) 37 to 54 (c) 11 to 18 (d) 55 to 86

23. Zero group or noble gases have general electronic configuration.
(a) ns^2, np^2 (b) ns^2, np^4 (c) ns^2, np^6 (d) ns^2, np^5

24. Which element exists in liquid form at room temperature?
(a) Ba (b) Br (c) Be (d) B

25. Number of elements placed in 6^{th} period is:
(a) 18 (b) 8 (c) 32 (d) 14

26. Point out the number of electrons in the valence shells of halogens:
(a) 6 (b) 5 (c) 7 (d) 8

27. Electronegativity of oxygen is:
(a) 2.5 (b) 3.0 (c) 3.4 (d) 4.0

28. Which one is the smallest among the following?
(a) Na (b) F (c) O (d) N

29. The radius of carbon atom is:
(a) 154 pm (b) 77 pm (c) 68 pm (d) 70 pm

30. Elements of group I and group II have valence electrons in:
(a) s-subshell (b) p-subshell (c) d-subshell (d) f-subshell

31. In 1860 correct atomic mass of elements were determined by:
(a) Cannizzaro (b) Newlands (c) Mosely (d) Mendeleev

32. How many elements were arranged by Mendeleev in order of increasing atomic masses?
(a) 60 (b) 61 (c) 62 (d) 63

33. How many periods are there in the modern periodic table?
(a) 6 (b) 7 (c) 8 (d) 9

34. 32 elements are present in which period of periodic table?
(a) 5^{th} (b) 6^{th} (c) 4^{th} (d) 8^{th}

35. Mark the incorrect statement.
(a) 1^{st} period contains two elements (b) 2^{nd} period contains eight elements
(c) 3^{rd} period contains 18 elements (d) 7^{th} period contains 23 elements only

36. In which atom outermost electrons are highly shielded?
(a) F (b) Cl (c) Br (d) I

37. Which of the following does not belong to 1st group?

(a) Hydrogen (b) Sodium (c) Rubidium (d) Calcium

38. Which one of the following halogen has highest electronegativity?

(a) F (b) I (c) Cr (d) CI

39. Point out among the following which has highest value of electron affinity

(a) F (b) CI (c) Br (d) I

40. Lanthanide series start after:

(a) La (b) Ba (c) Ra (d) Cs

41. ns^2 is the general electronic configuration of:

(a) Boron family (b) Nitrogen family
(c) Alkali metals (d) Alkaline earth metals

42. For Boron Z = 5, it belongs to which block:

(a) s (b) p (c) d (d) f

ANSWER KEY

1	c	12	a	23	c	34	b
2	b	13	b	24	b	35	d
3	c	14	b	25	c	36	d
4	a	15	b	26	c	37	d
5	b	16	a	27	c	38	a
6	c	17	c	28	b	39	a
7	d	18	c	29	b	40	a
8	d	19	c	30	a	41	d
9	a	20	a	31	a	42	b
10	d	21	b	32	d		
11	b	22	b	33	b		

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 4: Structure of Molecules
Exercise Questions

Exercise Multiple Choice Question Answer

1. **Atoms react with each other because:**
   (a) They are attracted to each other (b) They are short of electrons
   (c) They want to attain stability (d) They want to disperse
2. **An atom having six electrons in its valence shell will achieve noble gas electronic configuration by:**
   (a) Gaining one electron (b) Losing all electrons
   (c) Gaining two electrons (d) Losing two electrons
1. **Considering the electronic configuration of atoms which atom with the given atomic number will be the most stable one?**
   (a) 6 (b) 8 (c) 10 (d) 12
3. **Octet rule is:**
   (a) Description of eight electrons (b) Picture of electronic configuration
   (c) Pattern of electronic configuration (d) Attaining of eight electrons
4. **Transfer of electrons between atoms results in:**
   (a) Metallic bonding (b) Ionic bonding
   (c) Covalent bonding (d) Coordinate covalent bonding
1. **When an electronegative element combines with electropositive element the type of bonding is:**
   (a) Covalent (b) Ionic
   (c) Polar covalent (d) Coordinate covalent
5. **A bond formed between two non-metals is expected to be:**
   (a) Covalent (b) ionic (c) Coordinate covalent (d) Metallic
6. **A bond pair in covalent molecules usually has:**
   (a) One electron (b) two electrons (c) Three electrons (d) Four electrons
7. **Which of the following compounds is not directional in its bonding?**
   (a) C (b) KBr (c) CO_2 (d) H_2O
8. **Ice floats on water because:**
   (a) Ice is denser than water (c) Water is denser than ice
   (b) Ice is crystalline in nature (d) Water molecules move randomly
9. **Covalent bond involves the**
   (a) Donation of electrons (b) acceptance of electrons
   (c) Sharing of electrons (d) repulsion of electrons
10. **How many covalent bonds does C_2H_2 molecule have?**
    (a) Two (b) Three (c) Four (d) Five
11. **Triple covalent bond involves how many number of electrons?**

(a) Eight (b) Six (c) Four (d) Only three

12. Which pair of the molecules has same type of covalent bonds?

(a) O_2 and HCl (b) O_2 and N_2 (c) O_2 and C_2 (d) O_2 and C_2H_2

13. Identify the compound which is not soluble in water.

(a) C_6H_6 (b) NaCl (c) KBr (d) $MgCl_2$

14. Which one of the following is an electron deficient molecule?

(a) NH_3 (b) BF_3 (c) N_2 (d) O_2

15. Identify which pair has polar covalent bonds.

(a) O_2 and Cl_2 (b) H_2O and N_2 (c) H_2O and C_2H_2 (d) H_2O and HCl

16. Which one of the following is the weakest force among the atoms?

(a) ionic force (b) metallic force (c) intermolecular force (d) covalent force

ANSWR KEY

1	c	4	d	7	a	10	c	13	b	16	b
2	c	5	b	8	b	11	c	14	d	17	d
3	c	6	b	9	b	12	d	15	a	18	c

Exercise Short Question Answers

Q.1 Why do atoms react?

Ans: Atoms react to form chemical bonds in order to get stability. Atoms achieve stability by attaining electronic configuration of inert gases by losing, gaining or sharing of electron.

Q.2 Why is the bond between an electropositive and an electronegative atom ionic in nature?

Ans: The bond between an electropositive and an electronegative atom is ionic in nature because electropositive atom due to low I.E. can lose electron easily and forms a positive ion whereas electronegative atom due to high electron affinity will accept that electron easily and forms a negative ion. In this way positive and negative ions are attracted by electrostatic force of attraction to form ionic bond.

Q.3 Ionic compounds are solids. Justify.

Ans: Ionic compounds are solids because they have strong electrostatic forces of attraction between positively and negatively charged ions which hold them in a three dimensional crystalline or solid form.

Example:

Potassium chloride (KCl) is a crystalline solid.

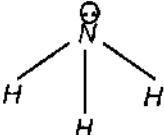
Q.4 More electronegative elements can form bonds between themselves. Justify.

Ans: More electronegative elements have high values of ionization energy and do not lose electrons. They share electrons between their own atoms to complete their valence shells and form covalent bond.

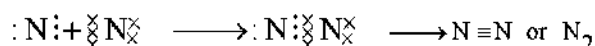
Q.5 Metals are good conductor of electricity. Why?

Ans: Metals are good conductors of electricity due to presence of mobile or free electrons.

Q.6 Ionic compounds conduct electricity in solution or molten form. Why?

Ans: Ionic compounds conduct electricity in solution or molten form because in these two states ionic compounds have free ions in them. When these free ions move in solution or molten state they become conductor of electricity.

Q.7 What type of covalent bond is formed in nitrogen molecule?

Ans: Triple covalent bond is formed in nitrogen molecule. In nitrogen molecule three bond pairs are involved in bond formation.

$$:\dot{N}\dot{:} + \overset{\times}{\underset{\times}{\times}}N\overset{\times}{\times} \longrightarrow :N\vdots\overset{\times}{\underset{\times}{\times}}N\overset{\times}{\times} \longrightarrow N\equiv N \text{ or } N_2$$

Q.8 Differentiate between lone pair and bond pair of electron.

Ans.

Bonded pair	Lone pair
i. Bond pair of electrons is involve in bond formation	**i.** Lone pair of electron is not involved in bond formation.
ii. Electrons of bond pair are contributed by two atoms.	**ii.** Electrons of lone pair are contributed by one atom only.
iii. It is under the influence of two nuclei (atoms)	**iii.** It is under the influence of only one nucleus.
In a ammonia molecule there are three bond pairs of electrons (N with lone pair, bonded to H, H, H)	

Q.9 Describe at least two necessary conditions for the formation of a covalent bond.

Ans: **Necessary conditions:**

- **a.** Elements should be electronegative in nature.
- **b.** Electronegativity difference between bonding atoms should be very small or zero.
- **c.** The elements should share the electrons mutually.

d. There should be 4 or more valance electrons.

e. The ionization energies of the elements must be high.

Example: HCl, Cl_2, C_6H_6 and C_2H_2

Q.10 Why HCl has dipole-dipole forces of attraction?

Ans: HCl forms a polar covalent bond atoms due to difference of electro negativity between bonded atoms. There exists a dipole in the molecule. The positive end of one molecule attracts the negative end of there molecule. Hence dipole force. (Intermolecular forces) exist between HCl molecules.

Example:

$$H^{\delta+} — Cl^{\delta-} \text{------} H^{\delta+} — Cl^{\delta-}$$

Q.11 What is a triple covalent bond, explain with an example?

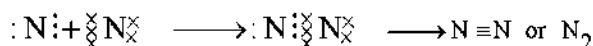
Ans: When each bonded atom contributes three electrons, three bond pairs are involved in bond formation. This type of bond is called triple covalent bond.

Representation:

It is represented by $(\equiv)$.

Example:

Triple covalent bond is formed in nitrogen molecule. In nitrogen molecule three bond pairs are involved in bond formation.

$$:\dot{N}\dot{:} + \overset{\times}{\underset{\times}{\times}}N\overset{\times}{\times} \longrightarrow :N\dot{:}\overset{\times}{\underset{\times}{\times}}N\overset{\times}{\times} \longrightarrow N \equiv N \text{ or } N_2$$

Q.12 What is difference between polar and non-polar covalent bonds, explain with one example of each?

Ans: Difference between polar and non polar covalent

Polar Covalent Bond	Non Polar Covalent Bond
i. It is a bond formed between two different types of atoms (hetero atoms).	*i. It is a bond formed between two similar atoms (homo atoms).*
ii. The shared pair of electrons is attracted by both the atoms un equally.	**ii.** The shared pair of electrons is attracted by both the atoms equally.
Examples. HCl, HBr, HF, H_2O etc	Examples. H_2, Cl_2, N_2, O_2 etc

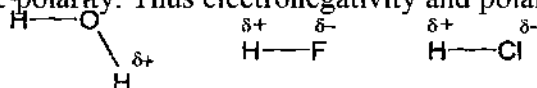
Q.13 Why a covalent bond becomes polar?

Ans: When there is a difference of electronegativity between two covalently bonded atoms, there will be unequal attraction for the bond pair of electrons between such atoms. It will result in the formation of polar covalent bond.

Examples: HCl, H_2O etc.

Q.14 What is relationship between electronegativity and polarity?

Ans: The polarity of a covalent bond depends upon the electronegativity difference between the bonded atoms. Higher the electronegativity difference between bonded atoms, greater will be the polarity. Thus electronegativity and polarity are directly related.

Examples: $H—O$ (with $\delta+$ H), $\overset{\delta+}{H}—\overset{\delta-}{F}$ $\overset{\delta+}{H}—\overset{\delta-}{Cl}$

Q.15 Why does ice float on water?

Ans: Ice floats on water because density of ice ($0.917 g/cm^3$) is less than that of liquid water ($1.00 g/cm^2$) at 0°C.

Q.16 Give the characteristic properties of ionic compounds.

Ans: Characteristics properties of ionic compounds.

i. Ionic compound are mostly crystalline solids.

ii. Ionic compounds are good conductors in solution and in molten form due to presence of free ions in them.

iii. Ionic compounds have high melting and boiling points. For example NaCl has melting point 800^0C and boiling point 1413°C.

iv. Ionic compounds dissolve in polar solvents e.g. NaCl dissolves in water.

Q.17 What characteristic properties do the covalent compounds have?

Ans: Characteristic properties of covalent compounds:

i. Melting boiling points: They have usually low melting and boiling point.

ii. Electrical conductivity: They are usually bad conductors of electricity. Polar compounds are conductors in their solutions in polar solvents.

iii. Solubility: They are usually insoluble in water but soluble in non-aqueous solvents like benzene, ether, alcohol and acetone.

iv. Crystal formation: Bigger molecules with three dimensional bonding form covalent crystals which are very stable and hard. They have high melting and boiling points.

Exercise Long Question Answers

Q.1 What is an ionic bond? Discuss the formation of ionic bond between sodium and chlorine atoms?

Ans: See Q. No. **4** (Subjective Part, Long Questions Answers)

Q.2 How can you justify that bond strength in polar covalent compounds is comparable to that of ionic compound?

Ans: See Q. No. 7 (Subjective Part, Long Questions Answers)

Q.3 What type of covalent bonds are formed between hydrogen, oxygen and nitrogen? Explain their bonding with dot and cross model.

Ans: See Q. No. **7** (Subjective Part, Long Questions Answers)

Q.4 How a covalent bond develops ionic character in it? Explain.

Ans:

Q.5 Explain the types of covalent bonds with at least one example of each type.

Ans: See Q. No. **5** (Subjective Part, Long Questions Answers)

Q.6 How a coordinate covalent bond is formed? Explain with examples?

Ans: See Q. No. **6** (Subjective Part, Long Questions Answers)

Q.7 What is metallic bonds? Explain the metallic bonding with the help of a diagram.

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.8 Define hydrogen bonding. Explain that how these forces affect the physical properties of compounds.

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.9 What are intermolecular forces? Compare these forces with chemical bond forces with reference to HCl molecule?

Ans:

Q.10 What is a chemical bond and why do atoms form a chemical bond?

Ans: See Q. No. **1** (Subjective Part, Long Questions Answers)

Q.11 What is octet rule? Why do atoms always struggle to attaint be nearest noble gas electronic configuration?

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 4: Structure of Molecules
Long Questions

Q.1 What is a chemical bond and why do atoms form chemical bonds? (Ex. Q.10)

Ans: *A chemical bond is defined as force of attraction between atoms that holds them together in a substance.*

Example: A bond formed between H and Cl atoms in a molecule of HCl.

Why do atoms form chemical bonds? (Atoms form bands to get stability)

It is a universal rule that everything in this world tends to become more stable.

Atoms achieve stability by attaining electronic configuration of inert gases (He, Ne or Ar etc) i.e. ns^2 np^6 having 2 or 8 electrons in the valence shell is sign of stability.

Duplet rule: Attaining two electrons in the valence shell is called duplet rule.

Octet Rule: An atom having eight electrons in the valence shell is called octet rule.

Why noble gases are non reactive?

The noble gases do .have 2 or 8 electrons in their valence shells. It means all the noble gases have their valence shells completely filled. Their atoms do not have vacant space in their valence shell to accommodate extra electrons. Therefore, noble gases do not gain, lose or share electrons. That is why they are non-reactive.

Importance of the noble gas electronic configuration:

The importance of the noble gas electronic configuration lies in the fact that all other atoms try their best to have the noble gas electronic configuration. For this purpose atoms combine with one another, which is called chemical bonding. In other words, atoms form chemical bonds to achieve stability by acquiring inert gas electronic configuration.

Q.2 What is octet rule? Why do atoms always struggle to attain the nearest noble gas electronic configuration? (Ex. Q.11)

Ans: Octet Rule:

The attaining of 8 electrons configuration in the valence shell, either by sharing, by losing or by gaining electrons is called octet rule.

Examples:

All noble gases except helium follow octet rule.

Explanation:

Atoms always struggle to attain the nearest noble gas electronic configuration in order to become more stable.

An atom can accommodate 8 electrons in its valence shell in three ways:

i. By giving valence shell electrons (if they are less than four) to other atoms.

ii. By gaining, electrons from other atoms (if the valence shell has five or more electrons in it)
iii. By sharing valence electrons with other atoms.

It means every atom has a natural tendency to achieve 2 or 8 electrons in its valence shell. The atoms having less than 2 or 8 electrons in their valence shells are unstable.

How can we identify the way an atom reacts?

The position of an atom in the periodic table indicates its group number. The group number is assigned on the basis of valence shell electrons.

Examples:

i. Group 1 has only 1 electron in its valence shell.
ii. Group 17 has 7 electrons in its valence shell.
iii. Mode of reaction of an atom depends upon its number of valence shell electrons.

Bond Formation

If the bond formation is between ions, it is due to an electrostatic force between them. But if bond formation is between similar atoms or between the atoms that have comparable electro negativities, then the chemical bond formation is by 'sharing' of electrons. This sharing of electrons may be mutual or one sided.

Effect of attractive and repulsive forces on bond formation:

When two approaching atoms come closer, the attractive as well as repulsive forces become operative. The formation of a chemical bond is a result of net attractive forces which dominate. The energy of that system is lowered and molecule is formed. Otherwise if repulsive forces become dominant no chemical bond will be formed. In that case there will be increase in the energy of the system due to creation of repulsive forces.

Q.3 Name the types of chemical bonds.

Ans: Types of chemical bond:

Bonding electrons:

The valence electrons, which are involved in chemical bonding, are termed as bonding electrons. They usually reside in the incomplete or partially filled outermost shell of an atom.

There are four types of chemical bonds depending upon the way how valence electrons are involved in bonding.

i. Ionic Bond
ii. Covalent Bond
iii. Dative Covalent or Coordinate Covalent Bond
iv. Metallic Bond

Q.4 What is ionic bond? Discuss the formation of ionic bond between sodium and chlorine atoms.

Ans: Ionic Bond:

The type of Chemical bond which is formed due to complete transfer of electron from one atom to another atom is called ionic bond.

OR

The electrostatic force of attraction between positive and negative ions is called ionic bond.

Explanation:

If the difference of electronegativity between two elements is more than 1.7 the bond between them will be permanently ionic bond.

The elements of Group-l and Group-2 being metals have the tendency to lose their valence electrons forming positively charged ions whereas non-metals of Group-15 to Group-17 have tendency to gain or accept electrons. They are electronegative elements with high electron affinities. If atoms belonging to these two different groups, metals and non-metals, are allowed to react and chemical bond is formed.

Conditions of ionic bond formation:

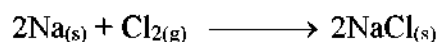
i. It is to be noted that only valence shell electrons take part in this type of bonding while other electrons are not involved.

ii. In such type of reaction heat is usually given out.

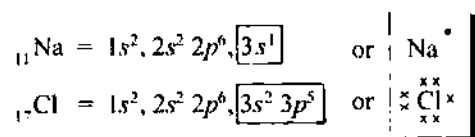
iii. The compounds formed due to this type of bonding are called ionic compounds.

Example: (Formation of sodium chloride, NaCl)

The formation of NaCl is a good example of this type of bond.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

Sodium chloride is a simple compound formed from Sodium (Z=11) and Chlorine (Z= 17). The ground state electronic configuration of these elements is shown below:

$$_{11}Na = 1s^2, 2s^2\ 2p^6, \boxed{3s^1} \quad \text{or} \quad Na^{\bullet}$$

$$_{17}Cl = 1s^2, 2s^2\ 2p^6, \boxed{3s^2\ 3p^5} \quad \text{or} \quad \overset{\times\times}{\underset{\times\times}{\stackrel{\times}{\times}Cl\times}}$$

The frame indicates electrons in valence shells of the elements; sodium has only one electron and chlorine has seven electrons. Sodium being electropositive element has the tendency to lose electron and chlorine being an electronegative element, has the tendency to gain electron. Therefore, they form positive and negative ions by losing and gaining electrons respectively, and their electronic configuration resembles with the nearest noble gases. $1S^2, 2S^2, 2p^6, 3s^2, 3p^6$ (Ar)

Formation of Na^+ ion:

Sodium atom loses one electron from the outermost shell and becomes sodium (Na^+) ion. Now the second shell becomes valance shell with 8 electrons.

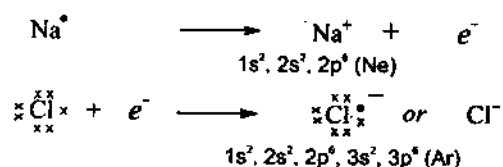
Formation of Cl^- ion:

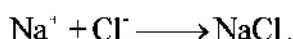
Chlorine atom gains one electron in the outermost shell and become Cl^- ion with 8 electrons.

$$Na^{\bullet} \longrightarrow Na^+ + e^-$$

$1s^2, 2s^2, 2p^6$ (Ne)

$$\overset{\times\times}{\underset{\times\times}{\stackrel{\times}{\times}Cl\times}} + e^- \longrightarrow \overset{\times\times}{\underset{\times\times}{\stackrel{\times}{\times}Cl\bullet}}^{-} \quad or \quad Cl^-$$

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ (Ar)

Na^+ and Cl^- ions stabilize themselves by combining with each other due to electro static force of attraction between them.

$$Na^+ + Cl^- \longrightarrow NaCl.$$

Q.5 Define the covalent bond. Explain the types of covalent bond.

OR

Explain the types of covalent bond with at least one example of each. (Ex. Q.5)

Ans: Covalent Bond:

"The type of bond, which is formed due to mutual sharing of electrons, is called covalent bond."

Explanation:

When bonding atoms have comparable values of electronegativity they share their electrons and form covalent bonds.

The elements of Group-13 to Group-I when allowed to react with each other, they form a chemical bond by mutual sharing of their valence shell electrons.

Formation of covalent bond (Energy changes during bond formation):

The energy changes during the covalent bond formation are of considerable value. When two atoms approach each other attractive forces develop between electrons of one atom and nucleus of other atom. Simultaneously repulsive forces between electrons of the two atoms as well as between their nuclei are also created. When the attractive forces dominate due to decrease in distance between those two atoms, a chemical bond is formed between them. By this mutual sharing of valence shell electrons each of the contributing atom attains the 'Octet' or nearest inert gas configuration

Bond pair:

The covalent bond is formed by mutual sharing of electrons between two atoms. The electrons that pair up to form a chemical bond are called 'bond pair' electrons.

Examples: Hydrogen chlorine, Nitrogen, Oxygen gases etc.

Types of covalent bonds:

Depending upon the number of bond pairs, covalent bond is classified into following three types:

i. Single Covalent bond.

ii. Double Covalent bond.

iii. Triple Covalent bond

i. Single Covalent Bond (–)

When one electron is contributed by each bonded atom, one bond pair is formed and it forms a single covalent bond.

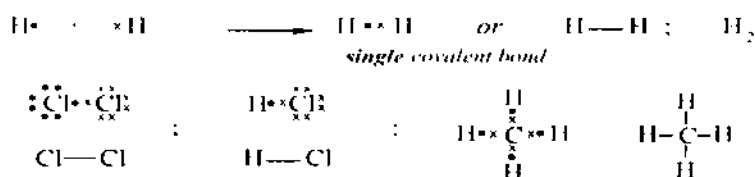
Representation:

It is indicated by single line (–) between two atoms.

Example: Hydrogen (H_2), Chlorine (Cl_2), Hydrochloric acid (HCl) and Methane (CH_4).

H• + ×H ⟶ H•×H *or* H—H ; H_2

single covalent bond

Cl—Cl ; H—Cl ; H—C—H (with H above and below C)

ii. Double Covalent Bond:

When two electrons are contributed by each bonded atom two bond pairs are formed and it forms a double covalent bond.

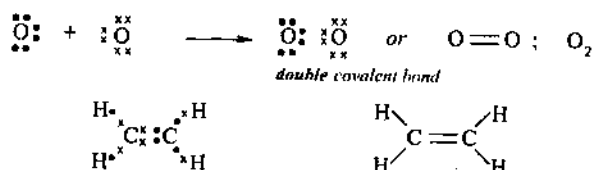
Representation:

It is indicated by two lines (=) between two bonded atoms.

Example: Oxygen (O_2) gas, Ethane (C_2H_4).

$$:\ddot{O}: + \,^{xx}_{xx}O \longrightarrow \ddot{O}::O \quad or \quad O = O\ ;\ O_2$$

double covalent bond

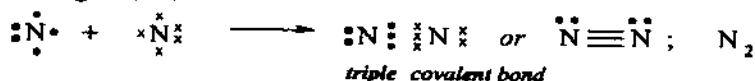
$$H\!:\!C::C\!:\!H \qquad H_2C = CH_2$$

iii. Triple Covalent Bond (≡)

When three electrons are contributed by each bonded atom, three bond pairs are formed and it forms a triple covalent bond.

Representation:

It is indicated by three lines (≡) between two bonded atoms.

Examples:

a. Nitrogen (N_2)

$$:\dot{N}\cdot + \,_x\dot{N}_x^x \longrightarrow :N\vdots\vdots N: \quad or \quad \ddot{N} \equiv \ddot{N}\ ;\ N_2$$

triple covalent bond

b. Ethyne (C_2H_2)

$$H\cdot\!\times C\vdots\vdots C\cdot\!\times H \qquad H - C \equiv C - H$$

By this mutual sharing of valence shell electrons, each of the contributing atom attains the 'octet' or nearest inert gas configuration.

Q.6 Define the coordinate covalent bond. Explain coordinate covalent bond with the help of example.

Ans: Dative covalent or coordinate covalent bond:

"Coordinate covalent or dative covalent bonding is a type of, covalent bonding in which the bond pair of electrons is donated by one bonded atom only."

Donor:

An atom which donates the electron pair is called donor

Acceptor:

An atom which accepts the electron pair is called acceptor.

Representation:

A small arrow (→) is usually used to indicate the atom and pair of electron being donated. The head of arrow is towards the acceptor atom

Lone pair of electrons:

The non bonded electron pair available on an atom in a molecule is called lone pair of electrons.

Example:

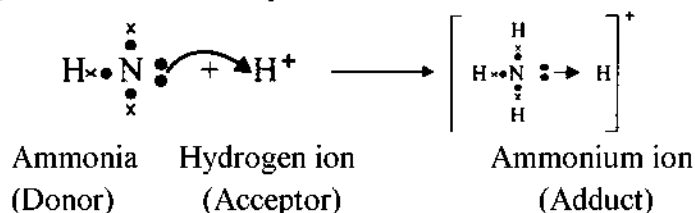
The electron pair available on nitrogen atom in ammonia (NH_3) molecule is called lone pair of electrons.

Examples:

i. Formation of ammonium radical (NH_4^{+1}):

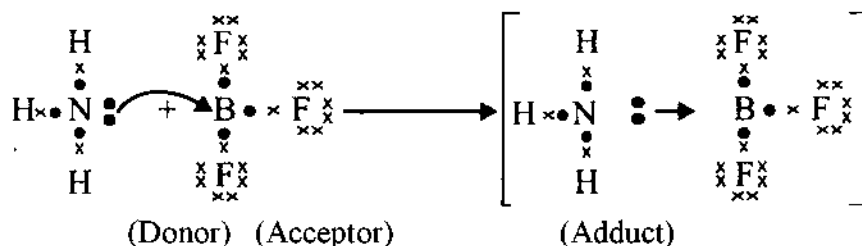
In the formation of ammonium ion, the nitrogen of NH_3 is the donor atom while hydrogen ion H^+ is the acceptor atom.

$$H \times\bullet \overset{\times}{\underset{\times}{N}} : + H^+ \longrightarrow \left[H \times\bullet \overset{H}{\underset{H}{N}} : \rightarrow H \right]^+$$

Ammonia (Donor)   Hydrogen ion (Acceptor)   Ammonium ion (Adduct)

ii. Formation of co-ordinate covalent bond between ammonia and boron tri-fluoride:

In the formation of BF_3 (boron tri fluoride) molecule, three valence electrons of boron atom ($Z = 5$) pair up with three electrons, one from each three fluorine atoms. The boron atom even after this sharing of electrons (covalent bond formation), remains short or deficient of two electrons in its outermost shell. Now if a molecule with a lone pair approaches this molecule, it accepts lone pair from that donor and forms a coordinate covalent bond. The one pair on nitrogen of ammonia molecule makes it a good donor molecule to form a coordinate covalent bond.

$$H \times\bullet \overset{H}{\underset{H}{N}} : + \overset{F}{\underset{F}{B}} \bullet \times F \longrightarrow \left[H \times\bullet \overset{H}{\underset{H}{N}} : \rightarrow \overset{F}{\underset{F}{B}} \bullet \times F \right]$$

(Donor)   (Acceptor)   (Adduct)

Q.7 Explain in detail the Polar and non polar Covalent bond.

Ans: Polar and Non-polar Covalent Bond.

i. Non-Polar Covalent Bond:

A covalent bond formed between two similar atoms (homo-atoms) in shared pair of electrons is attracted by both the atoms equally, called non-polar covalent bond.

Explanation:

These bonds are formed by equal sharing of electron pair between the two bonding atoms. This type of bond is called a pure covalent bond.

Examples: Bond formation in H_2, Cl_2, O_2, N_2 and F_2

ii. Polar Covalent Bond

A covalent bond formed between two atoms of different elements (hetero-atom) in which shared pair of electrons is attracted by both the atoms unequally is called polar covalent bond.

Examples: Water, hydrogen fluoride, hydrogen chloride etc.

Formation of polar bond:

The difference between electro negativities of hydrogen and chlorine is 1.0. As the electro-negativity of chlorine is more, it attracts the shared pair of electron towards itself with a greater force. A partial negative charge is therefore created on chlorine and

in turn a partial positive charge on hydrogen due to electro negativity difference. It creates polarity in the bond and is called polar covalent bond.

Delta (δ) sign:

The delta (δ) sign indicates partial positive or partial negative charge that is developed due to unequal sharing of shared pair or bonded pair of electrons.

The compounds resulting from polar covalent bonds are called polar compounds.

Determination of Nature of Chemical bond:

By using electro negativity values, it is possible to predict whether a chemical bond will be ionic or covalent in nature.' A bond formed between elements of high electro negativity (halogen group) and elements of low electro negativity (alkali metals) are ionic in nature there is complete transfer of electrons between them.

If the difference of electro negativities between two elements is more than 1.7 the bond between them will be predominantly ionic bond and if it is less than 1.7, the bond between two atoms will be predominantly covalent.

Q.8 What is metallic bond? Explain metallic bonding with the help of diagram. (Ex. Q7)

Ans: Metallic Bond:

The metallic bond is defined as a bond formed between metal atoms (positively charged ions) due to mobile or free electrons.

Properties:

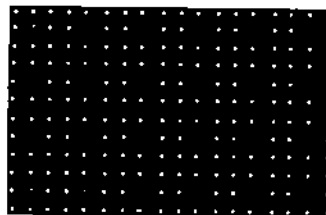
i. They have high melting and boiling points.

ii. They show good conductions of heat and electricity.

iii. They are hard and of heavy nature.

Reason for the formation of metallic bond in metals:

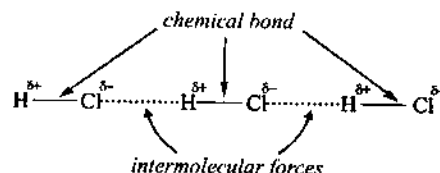
i. In case of metals, the hold of nucleus over the outermost electrons is weak because of large sized atoms and greater number of shells in between nucleus and valence electrons.

ii. Because of low ionization potentials, metals have the tendency to lose their outer electrons easily. The loose 'or free electrons of all metal atoms move freely in the spaces between atoms of a metal. None of these electrons is attached to any particular atom. They belong to a common pool or belong to all the atoms of that metal. Nuclei of metal atoms appear submerged in sea of these free mobile electrons. The mobile electrons are responsible for holding the atoms of metals together forming a metallic bond.

Example: The bond found in Na, Ca, Mg etc.

Q.9 What are intermolecular forces? Explain their types in detail.

Ans: Intermolecular forces:

"*The forces of attraction presence between molecule of a substance are called inter molecular forces.:*

The forces that hold atoms in a compound are chemical bonds. In addition to these strong bonding forces, relatively weak forces also exist in between the molecules, which are called intermolecular forces.

Comparison of strength of intermolecular and intermolecular forces:

It requires about 17 kJ energy to break these intermolecular forces between one mole of liquid hydrogen chloride molecules to convert it into gas whereas, about 430 kJ are required to break the chemical bond between hydrogen and chlorine atoms in 1 mole of hydrogen chloride.

Types of Inter molecular forces:

All intermolecular forces, which are collectively called van der Waals forces, are electrical in nature. Following are types of intermolecular forces

i. Dipole-Dipole Forces

ii. Hydrogen Bonding

iii. Dipole - Induced Dipole Forces

iv. Ion Dipole Forces

v. London Forces

i. Dipole - Dipole Interaction:

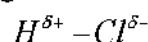
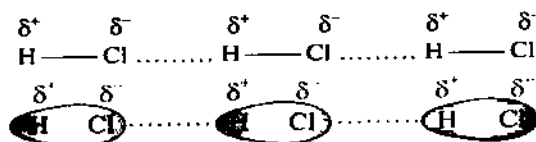
The force of attraction present between partial positive end of one polar molecule and partial negative and of other polar molecule is called dipole - dipole force.

Explanation:

i. They result from the attractions of opposite charges which may be temporary or permanent.

ii. The unequal sharing of electrons between two different types of atoms make one end of molecule slightly positive and other end slightly negatively charged.

iii. As shared pair of electron is drawn towards more electronegative atom, it is partially negatively charged, as chlorine in hydrogen chloride. The other end automatically becomes partially positively charged.

$$H^{\delta+} - Cl^{\delta-}$$

iv. When partial positive and partial negative charges exist at different positions in a molecule, the adjacent molecules will arrange themselves in such a way that negative portion of that molecule comes near to positive portion of other molecule.

ii. Hydrogen Bonding:

The forces of attraction present between partially positive hydrogen atom of one molecule and highly electronegative atom (N, O or F) of another molecule is called hydrogen bonding.

Partially positively hydrogen of one molecule attracts and forms a bond with the partially negatively charge atom of the other molecule, the bonding is called hydrogen bonding

Explanation:

Occurrence: Hydrogen bonding is a special type of intermolecular forces present in the permanently polar molecules. This bonding can be considered unique dipole-dipole attraction.

Development of hydrogen bonding: This force of attraction develops between molecules that have a hydrogen atom bonded to a small, highly electronegative atom with lone pairs-of electrons such as nitrogen, oxygen and fluorine. The covalent bond between hydrogen atom and other atom becomes polar enough to create a partial positive charge on hydrogen atom and a partial negative charge on the other atom. The small size and high partial positive charge on the hydrogen atom enables it to attract highly electronegative (N, O or F) atom of the other molecule.

Representation:

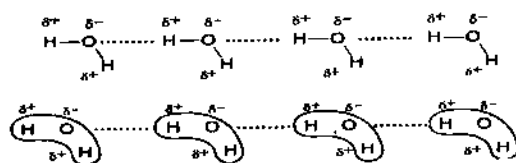
This force of attraction is represented by a dotted line (……) between the molecules as shown below:

Hydrogen bonding and physical properties:

a. Boiling Points:

i. Due to this, boiling points of the compounds are affected greatly.

ii. It enhances the force of attraction between molecules.

Boiling point of water (100°C) is higher than that of alcohol (78°C) because of more and stronger hydrogen bonding in water.

b. Floating of Ice:

The important phenomenon of floating of ice over water is because of hydrogen bonding. The density of ice at 0 °C (0.917 gcm^{-3}) is less than that of liquid water at 0°C (1.00 g/cm^3). In the liquid state water molecules move randomly, however. When water freezes the molecules arrange themselves in an ordered form that gives them open structure. This process expands the molecules. That results in ice being less dense as compared to water.

4.5 Nature Of Bonding and Properties:

Q.10 Write down the properties of ionic compounds.

Ans: Ionic Compounds:

"The compounds which contain ionic bond in them are called ionic compounds."

Examples: NaCl, KCl, KNO_3, $CaCO_3$ etc.

Composition:

Ionic compounds are made up of positively and negatively charged ions. Thus they consist of ions and not the molecules.

Attractive forces:

These positively and negatively charged ions are held together in a solid or crystal form with strong electrostatic attractive forces.

Order arrangement of ions:

The orderly arrangement of Na+ and Cl^{-1} ions in a solid crystal of sodium chloride is given below.

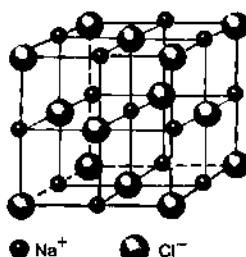



Figure 4.3 Regular arrangement of Na^+ and Cl^- ions in solid crystal of NaCl

Properties of Ionic Compounds:

The ionic compounds have following properties:

i. Crystalline Solids: Ionic compounds are mostly crystalline solids.

ii. Electrical conductivity: Ionic compounds in solid state have negligible electrical conductance but they are good conductors in solution and in the molten form. It is due to presence of free ions in them.

iii. Melting and boiling points: Ionic compounds have high melting and boiling points. **For example,** sodium chloride has melting point 800°C and a boiling point 1413 °C. As ionic compounds are made up of positive and negative ions, there exist strong electrostatic forces of attraction between oppositely charged ions. So, a great amount of energy is required to break these forces.

v. **Solubility:** They dissolve easily in polar solvents like water. Water has high dielectric constant that weakens the attraction between ions.

Q.11 What are covalent compounds?

Ans: Covalent Compounds:

"The compounds which contain covalent bond in them are called covalent compounds."

Composition:

The covalent compounds are made up of molecules that are formed by sharing of

electrons between their atoms.

Strength of bond:

A covalent bond is generally regarded as weaker than an ionic bond. Covalent compounds are made up of two or more non-metals.

Physical states:

Lower molecular mass covalent compounds are gases or low boiling liquids. Contrary to it, higher molecular mass covalent compounds are solids.

Example: H_2, Cl_2, CO_2, H_2SO_4, $C_6H_{12}O_6$ etc.

Properties of Covalent Compounds:

i. Melting and boiling points: They have usually low melting and boiling points.

ii. Electrical conductivity: They are usually bad conductors of electricity. The compounds having polar character in their bonding are conductor of electricity when they dissolve in polar solvents.

iii. Solubility: They are usually insoluble in water but are soluble in non-aqueous solvents like benzene, ether, alcohol and acetone.

iv. Crystal formation: Bigger molecules with three dimensional bonding form covalent crystals which are very stable and hard. They have high melting and boiling points.

Q.12 Write down the properties of polar and non-polar compounds.

Ans: a. Polar Compounds:

"A compound having polar molecule is called polar compound."

Examples: HF, HCl, H_2O, NH_3 etc

Development of polarity in chemical:

Polarity in a chemical bond is due to difference in electro negativities of the bonding atoms.

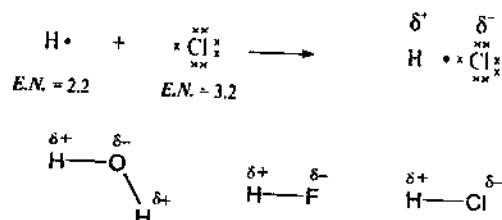
Scale:

On the Pauling Scale fluorine has been given an electronegativity value of 4. The values for other elements are calculated relative to it.

Properties:

i. Properties of non-polar and polar covalent compounds differ to some extent.

ii. Polar covalent compounds usually dissolve in water while non polar do not dissolve.

iii. An aqueous solution of a polar compound usually conducts electricity due to the formation of ions as a result of its reaction with water.

b. Non-Polar Compounds:

"A compound having non polar molecule is called non polar compound."

Examples: CO_2, CH_4, C_6H_6, C_2H_2, CCl_4 etc.

Properties:

i. Non-polar covalent compounds usually do not dissolve in water
ii. Similarly non-polar compounds do not conduct electricity

Q.13 Write down the properties of coordinate covalent compounds.

Ans: Coordinate Covalent Compounds:

The compounds which contain coordinate covalent bond in them are called coordinate covalent compounds.

Examples: $NH_3 - BF_3$, NH_4Cl , NH_3AlCl_3 etc.

Properties:

i. Their properties are mostly similar to those of covalent compounds.
ii. As the nuclei in these compounds are held by shared electrons, therefore, they do not form ions in water.
iii. Due to their covalent nature they form solutions in organic solvents and are very less soluble in water.
iv. Usually they are rigid compounds with a dipole.

Q.14 Write down the properties of metals.

Ans: Metals:

"The elements which are usually hard, are good conductors of heat and electricity and are malleable and ductile are called metals."

Metals have common property of conducting heat and electricity. It gives them prime role in many technologies.

Examples: Iron, cobalt, nickel, gold, silver etc.

Properties:

i. They show **metallic luster.**
ii. They are usually **malleable and ductile.** Malleability is the property by virtue of which a metal can be drawn into sheets, while ductility is the property by virtue of which a metal can be drawn into wires.
iii. They have usually **high melting and boiling** points.
iv. Being greater in size they have **low ionization energies and form cations** (M) very easily.
v. They are **good conductors of heat and electricity** in solid and liquid state due to mobile electrons
vi. Metals have **shining surface.**

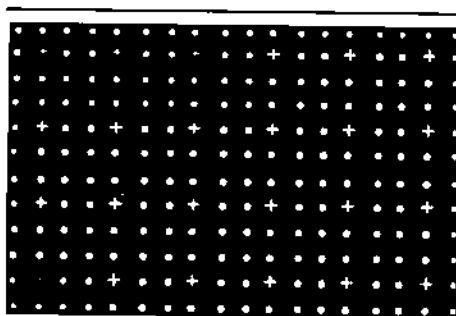



Fig. 4.2 A schematic diagram of Copper wire showing its positive nuclei (+) embedded in sea of free electrons (o) making 'Metallic Bonding'

Q.15 What are synthetic adhesives? Describe properties of epoxy adhesives.

Ans: Synthetic Adhesives:

"The synthetically produced substances which are used to stick the surfaces are called synthetic adhesives"

Properties:

i. Although natural adhesives are less expensive to produce,

ii. But most important adhesives used nowadays are synthetic.

iii. Adhesives based on synthetic resins and rubbers excel in versatility and performance.

iv. Synthetics adhesives can be produced in a constant supply with uniform properties and they can be modified in many ways.

v. The polymers or resins used in synthetic adhesives fall into two general categories-thermoplastics and thermosetting.

Example: One form of polymer used industrially is epoxy adhesive.

Epoxy adhesive:

"Epoxy is polymer that is formed from two different chemicals. These are referred to as resin and the hardener. Epoxy adhesives are called structural adhesives."

Uses:

i. These high-performance adhesives are used in the construction of aircraft, automobiles, bicycles, boats, golf clubs, where high strength bonds are required.

ii. Epoxy adhesives can be developed to suit almost any application.

Properties:

i. They can be made flexible or rigid, transparent or opaque even colored as well as fast or slow setting.

ii. Epoxy adhesives are good heat and chemical resistant

iii. They are stable to heat up to temperatures 177°C.

Because of these properties, they are given the name of engineering adhesives.

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 4: Structure of Molecules
Extra MCQ's

1. **Elements of group 18 are called:**
   (a) Noble metals (b) Noble gases (c) Noble (d) None of these
2. **Element attaining eight electrons in its outermost shell are called**
   (a) Duplet rule (b) octet rule (c) Both "A" and "B" (d) None of these
3. **The force of attraction which holds the molecule atom together is called**
   (a) Ionic bond (b) Covalent bond (c) hydrogen bond (d) Chemical bond
4. **Every atom tries to attain**
   (a) Maximum energy (b) Stable energy (c) Minimum energy (d) Covalent energy
5. **Which force of attraction is present between positive and negative ions?**
   (a) Homolytic (b) Electrostatic (c) Electro elastic (d) Electronic
6. **A bond formed by the complete transfer of electron from one atom to other is called**
   (a) Chemical bond (b) Covalent bond (c) Ionic bond (d) metallic bond
7. **Hydrogen acquires the electronic configuration like**
   (a)Xe (b) Ne (c) He (d) Ar
8. **In non polar covalent bond the bonded atoms share the electrons**
   (a) Not equally (b) Equally (c) Differently (d) Oppositly
9. **In polar covalent bonds molecules are**
   (a) Homo atomic (b) Tri atomic (c) Hetero atomic (d) Mono atomic
10. **Dipole-dipole interactions are**
    (a) Magnetic (b) Electric (c) Neutral (d) Stable
11. **Energy requires to break forces of one mole liquid HCl molecule is**
    (a) 16kJ (b) 15kJ (c) 17kJ (d) 18kJ
12. **The density of ice at 0°C is**
    (a) 0.917 gcm^{-3} (b) 0.719 gcm^{-3} (c) 0.197 gcm^{-3} (d) 0.0917 gcm^{-3}
13. **One form of polymer used industrially as a adhesive**
    (a) Epoxy (b) Explosive (c) Synthetic (d) Natural
14. **Epoxy is a polymer formed by the resin and**
    (a) Shiner (b) Softener (c) Hardener (d) Thiner
15. **Substance have the tendency to gain electrons and anions**
    (a) Metals (b) Non-metals (c) Metalloids (d) Halogens
16. **Which properties are affected by hydrogen bonding**
    (a) Physical (b) Chemical (c) Ionic (d) Metallic
17. **Which scientist had given the idea of octaves in periodic table?**
    (a) Mendeleev (b) Al-razi (c) Newland (d) Dobereiner
18. **Covalent bond involves the**
    (a) Donation of electrons (b) Acceptanee of electrons
    (c) Sharing of electrons (d) Repulsion of electrons
19. **Triple covalent bond involves how many numbers of electrons?**
    (a) Eight (b) Six (c) Four (d) Only three
20. **Which one of the following is an electron deficient molecule?**

(a) NH_3 (b) BF_3
(c) N_2 (d) O_2

21. Which one of the following is the weakest force among the atoms?
(a) Ionic force (b) Metallic force
(c) Intermolecular force (d) Covalent force

22. Which atom after losing one electrons attains electronic configuration of noble gas.
(a) He (b) Ne
(c) Ar (d) Na

23. How many electrons are there in the outer most shell of Cl-atom after it gains one electron in its values shell?
(a) 6 (b) 7
(c) 8 (d) 10

24. Which ionic species has same electronic configuration as noble gas Ar.
(a) Na^+ (b) K^+
(c) Cl^- (d) Both (b) and (c)

25. An ionic bond is formed when elements of group 1 react with elements of group:
(a) 12 (b) 14
(c) 18 (d) 17

26. A dative bond is formed between ammonia and borontrifluoride the donar atom is:
(a) Fluorine (b) Boron
(c) Hydrogen (d) Nitrogen

27. Indicate which molecule is non-polar:
(a) NaCl (b) KBr
(c) CO_2 (d) Kl

28. Mark the molecule which possess four covalent bonds.
(a) N_2 (b) C_2H_4
(c) CH_4 (d) C_2H_2

29. Which one of molecule is maximum polar:
(a) HCl (b) HF
(c) NaF (d) H_2O

30. Electronegativity difference between hydrogen and chlorine is:
(a) 1.7 (b) 0.8
(c) 1.0 (d) 0.5

31. Indicate which of the following has hydrogen bonding:
(a) C_2H_6 (b) CH_4
(c) NH_3 (d) HCl

32. Is^2, $2s^2$, $2p^2$ is the electronic configuration of which atom.
(a) Li (b) B
(c) N (d) C

33. Identify the compound which is soluble in water:
(a) C_6H_6 (b) CH_4
(c) C_2H_4 (d) NaCl

34. The bond dissociation energy of H–Cl molecule
(a) 430 kj/mole (b) 340 kj/mole
(c) 403 kj/mole (d) 304 kj/mole

35. On the pauling scale fluorine has been given an electronegative value of:
(a) 2.5 (b) 3.0
(c) 3.5 (d) 4.0

36. The density of ice at 0°C is:
(a) $0.917 g\ cm^{-3}$ (b) $1.0\ gcm^{-3}$

(c) 0.719g cm^{-3} (d) 0.17g cm^{-3}

37. Which molecule contains a single covalent bond?

(a) CH_4 (b) C_2H_2 (c) C_2H_4 (d) O_2

38. The difference between elecronegativities of hydrogen and fluorine

(a) 1.0 (b) 3.0 (c) 2.0 (d) 1.8

39. The electronegativity of hydrogen atom is

(a) 2.0 (b) 2.2 (c) 3.0 (d) 2.1

40. The energy required to break the intermolecular forces between one mole of liquid hydrogen chloride molecule to convert into gas is

(a) 22kJ (b) 132 kJ (c) 32 kj (d) 17 kJ

41. the energy required to break the chemical bond between hydrogen and chlorine atoms in 1 mole of hydrogen chloride is

(a) 320 kJ (b) 430 kJ (c) 365 kJ (d) 410 kJ

42. The boiling point of water is

(a) 0°C (b) 35°c (c) 100°c (d) 25°c

43. The boiling point of alcohol is

(a) 44°C (b) 19°C (c) 53°C (d) 78°C

44. Epoxy adhesives are stable to heat up to a temperature of

(a) 177°C (b) 225°C (c) 320°C (d) 135°C

45. The density of water at 0°C is

(a) 2.0 g/cm^3 (b) 1.0 g/cm^3 (c) 0.917 g/cm^3 (d) 1.17 g/cm^3

46. The boiling point of NaCl is

(a) 318°C (b) 1413°C (c) 1215°C (d) 1510°C

47. Which of the following is an example of covalent compound?

(a) $C_6H_{12}O_6$ (b) CH_4 (c) H_2SO_4 (d) All of these

48. Every atom has a natural tendency to achieve electrons in its valence shell

(a) 2 or 6 (b) 2 or 8 (c) 2 or 4 (d) 2 or 10

49. The formation of ionic bond between two ions is due to

(a) Hydrogen bonding (b) Electrostatic forces

(c) Metallic forces (d) All of them

50. How many valence shell electrons are there in Na^+ ion?

(a) 8 (b) 9 (c) 10 (d) 11

ANSWER KEY

1	b	11	c	21	c	31	c	41	b
2	b	12	a	22	c	32	d	42	c
3	d	13	a	23	c	33	d	43	d
4	c	14	d	24	c	34	a	44	a
5	b	15	d	25	d	35	d	45	c
6	c	16	a	26	d	36	a	46	b
7	c	17	c	27	c	37	a	47	d
8	b	18	c	28	c	38	d	48	b
9	c	19	b	29	b	39	b	49	b
10	c	20	b	30	c	40	d	50	b

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 5: Physical States of Matter

Exercise Questions

Exercise Multiple Choice Question Answers

1. **How many times liquids are denser than gases?**
   (a) 100 times (b) 1000 times (c) 10,000 times (d) 100,000 times
2. **Gases are the lightest form of matter and their densities are expressed in terms of:**
   (a) mg cm^{-3} (b) g cm^{-3} (c) g dm^{-3} (d) kg dm^{-3}
3. **At freezing point which one of the following coexists in dynamic equilibrium:**
   (a) Gas and solid (b) liquid and gas (c) liquid and solid (d) all of these.
4. **Solid particles possess which one of the following motions?**
   (a) Rotational motions (b) vibrational motions
   (c) Translational motions (d) both translational and vibrational motions
5. **Which one of the following is not amorphous?**
   (a) Rubber (b) plastic (c) glass (d) glucose.
6. **One atmospheric pressure is equal to how many Pascals:**
   (a) 101325 (b) 10325 (c) 106075 (d) 10523
7. **In the evaporation process, liquid molecules which leave the surface of the liquid have:**
   (a) Very low energy (b) moderate energy (c) very high energy (d) none of these
8. **Which one of the following gas diffuses faster?**
   (a) Hydrogen (b) helium (c) fluorine (d) chlorine
9. **Which one of the following does not affect the boiling point?**
   (a) Intermolecular forces (b) external pressure
   (c) Nature of liquid (d) initial temperature of liquid
10. **Density of a gas increases, when its:**
    (a) Temperature is increased (b) pressure is increased
    (c) Volume is kept constant (d) none of these
11. **The vapour pressure of a liquid increases with the:**
    (a) Increase of pressure (b) increase of temperature
    (c) Increase of intermolecular forces (d) increase of polarity of molecules

ANSWR KEY

1	a	4	b	7	c	10	b
2	c	5	d	8	a	11	b
3	c	6	a	9	d		

Exercise Short Question Answers

Q.1 What is diffusion, explain with an example?

Ans: *The spontaneous mixing of particles of a substance by random motion and collisions, to form a homogeneous mixture is called diffusion.*

OR

Movement of molecules of a substance from the region of higher concentration to the region of lower concentration is called diffusion.

Example: When a few drops of ink are added in beaker of water, ink molecules move around and after a while spread in whole of the beaker. Thus diffusion has taken place.

Q.2 Define standard atmospheric pressure. What are its units? How it is related to Pascal?

Ans: **Standard atmospheric pressure:**

It is the pressure exerted by the atmosphere at the sea level. *"It is defined as **the pressure exerted by a mercury column of 760mm height at sea level"**. It is sufficient pressure to support a column of mercury 760mm in height at sea level.*

Units:

i. One atmosphere (1 atm): 1 atm is called standard pressure

ii. One pascal (1 Pa)

1atm = 760mmHg = 760torr = $101325Nm^{-2}$ = 101325Pa

(as 1mmHg = 1torr

$1Nm^{-2}$ = 1Pa)

Q.3 Why are the densities of gases lower than that of liquids?

Ans: Gases have lower densities than densities of liquids. It is due to the light mass and more volume occupied by the gases. Another reason for lower densities of gases is negligible intermolecular forces among the gases molecules. On the other hand liquid molecules are closely spaced and have strong intermolecular forces.

Q.4 What do you mean by evaporation, how it is affected by surface area?

Ans: **Evaporation:**

"*The process of changing of a liquid into a gas phase is called evaporation.*"

Affect of surface area on evaporation:

Evaporation is a surface phenomenon. Greater is surface area, greater is evaporation and vice versa.

Q.5 Define the term allotropy with examples.

Ans: Allotropy:

"The existence of an element in more than one forms, in same physical state is called allotropy."

Examples:

i. Oxygen has two allotropic forms i.e. oxygen (O_2) and ozone (O_3).

ii. Three allotropic forms of carbon are: Diamond, graphite and bucky balls.

Q.6 In which form sulphur exists at 100°C?

Ans: Sulphur exists in monoclinic form at 100°C

Q.7 What is the relationship between evaporation and boiling point of a liquid?

Ans: Relationship between evaporation and boiling point:

If the boiling point of a liquid is high, its evaporations slow. Because intermolecular forces are high in the liquid which have high boiling points. If boiling point is low then evaporation is high.

Exercise Long Question Answers

Q.1 Define Boyle's law and verify it with an example.

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Q.2 Define and explain Charles law of gases.

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Q.3 What is vapour pressure and how it is affected by intermolecular forces.

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.4 Define boiling point and also explain, how it is affected by different factors.

Ans: See Q. No. **10** (Subjective Part, Long Questions Answers)

Q.5 Describe the phenomenon of diffusion in liquids along with factors which influence it.

Ans: See Q. No. **12** (Subjective Part, Long Questions Answers)

Q.6 Differentiate between crystalline and amorphous solids.

Ans: See Q. No. **15** (Subjective Part, Long Questions Answers)

Exercise Solved Numericals

Q.1 <u>Convert the following units:</u>

a. 850 mm Hg to atm

b. 205000 Pa to atm

c. 560 torr to cm Hg

d. 1.25 atm to Pa

Solution:

a. 850 mmHg to atm

760mmHg = 1atm

1mmhg $= \frac{1}{760} atm$

850mmHg $= \frac{1}{760} \times 850 atm$

= 1.12atm

b. 205000 Pa to atm

101325Pa = 1atm

1Pa $= \frac{1}{101325} atm$

205000Pa $= \frac{1}{101325} \times 205000$ atm

= 2.02 atm

c. 560 torr to cm Hg

760 torr = 760 mmHg

= 76 cmHg

1 torr $= \frac{76}{760}$ cmHg

560 torr $= \frac{76}{760} \times 560 cmHg$

= 56cmHg

d. 1.25 atm to Pa

1 atm = 101325Pa

1.25 atm = 1.25×101325Pa

= 126656 Pa

Q.2 <u>Convert the following units.</u>

a. 750°C to K **b. 150°C to K** **c. 100 K to °C** **d. 172 K to °C**

Solution:

a. 750°C to K

T (°C) = 750 °C

T (K) = ?

T (K) = T(°C)+ 273

= 750 + 273

= 1023K

b. 150°C to K

$$T(°C) = 150\ °C$$
$$T(K) = ?$$
$$T(K) = T(°C) + 273$$
$$= 150 + 273$$
$$= 423$$

c. 100 K to °C

$$T(K) = 100\ K$$
$$T(°C) = ?$$
$$T(°C) = T(K) - 273.15$$
$$= 100 - 273$$
$$= -173°C$$

d. 172 K to °C

$$T(K) = 172\ K$$
$$T(°C) = ?$$
$$T(°C) = T(K) - 273$$
$$= 172 - 273$$
$$= -101°C$$

Q.3 <u>A gas at pressure 912 mm of Hg has volume 450cm³. What will be its volume at 0.4 atm.</u>

Given Data:

$$P_1 = 912\text{ mm Hg} = \frac{912\text{ mm Hg}}{760\text{ mm Hg}}$$
$$= 1.2\text{ atm}$$
$$V_1 = 450\text{ cm}^3$$
$$P_2 = 0.4\text{ atm}$$

Required:

$$V_2 = ?$$

Using the equation of Boyle's Law:

$$P_1V_1 = P_2V_2$$

Solution:

By putting the values:

$$1.2\text{ atm} \times 450\text{ cm}^3 = 0.4\text{ atm} \times V_2$$
$$V_2 = \frac{1.2\text{ atm} \times 450\text{ cm}^3}{04\text{ atm}}$$
$$V_2 = \frac{12}{4} \times 450\text{ cm}^3$$
$$V_2 = 3 \times 450\text{ cm}^3$$
$$\mathbf{V2 = 1350\ cm^3}$$

Q.4 A gas occupies a volume of 800 cm³ at 1 atm, when it is allowed to expand up to 1200 cm³ what will be its pressure in mm of Hg.

Given Data:

$P_1 = 1 \text{ atm}$

$V_1 = 800 \text{ cm}^3$

$V_2 = 1200 \text{ cm}^3$

Required:

$P_2 = ?$

Using the equation of Boyle's Law:

$P_1V_1 = P_2V_2$

Solution:

By putting the values

$$1 \text{ atm} \times 800 \text{ cm}^3 = P_2 \times 1200 \text{ cm}^3$$

$$P_2 = \frac{1 \text{ atm} \times 800 \text{ cm}^3}{1200 \text{ cm}^3}$$

$$P_2 = \frac{2}{3} \text{ cm}^3$$

$$P_2 = 0.667 \text{ atm}$$

As $1 \text{ atm} = 760\text{mmHg}$

So $0.66 \text{ atm} = 760 \times 0.66\text{mmHg}$

$= 506.66\text{mmHg}$

Q.5 It is desired to increase the volume of a fixed amount of gas from 87.5 to 118 cm³ while holding the pressure constant. What would be the final temperature if the "initial temperature is 23°C.

Given Data:

$V_1 = 87.5 \text{ cm}^3$

$V_2 = 118 \text{ cm}^3$

$T_1 = 23°C\ (23+273)\ K = 296K$

Required:

$T_2 = ?$

By using the equation of charle's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Solution

$$T_2V_1 = V_2 \times T_1$$

Or

$$T_2 = \frac{V_2 \times T_1}{V_1}$$

By putting the values

$$T_2 = \frac{118\text{cm}^3 \times 296K}{87.5 \text{ cm}^3}$$

$$T_2 = 399K$$

T_2 can be converted into Celsius scale as:

$$T_2 = 299 - 273 = 126°C$$

Q.6 **A sample of gas is cooled at constant pressure from 30°C to 10°C. Comment:**

a. **Will the volume of the gas decrease to one third of its original volume?**

b. **If not, then by what ratio will the volume decrease?**

Solution:

a.

$T_1 = 30°C = (30+273)\ K = 303K$

$T_2 = 10°C = (10+273)K = 283K$

$V_1 = 1\ dm^3$

Required: $V_2 = ?$

Solution:

By using the equation of Charle's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = \frac{V_1}{T} \times T_2$$

By putting the values

$$= \frac{1dm^3}{303K} \times 283K$$

$$V_2 = 0.93dm^3$$

The volume of gas will not decrease to one third of its original volume.

(b) The volume decreases in the ration 1:0.93.

Q.7 **A balloon that contains 1.6 dm³ of air at standard temperature and pressure is taken under water to a depth at which its pressure increases to 3.0 atm. Suppose that temperature remain unchanged, what would be the new volume of the balloon. Does it contract or expand?**

Given Data:

$P_1 = 1\ atm$

$V_1 = 1.6\ dm^3$

$P_2 = 3.0\ atm$

Required:

$V_2 = ?$

Solution:

By using the equation of Boyle's law

$$P_1V_1 = P_2V_2$$

By putting the values

$$1\ atm \times 1.6\ dm = 3\ atm \times V_2$$

$$V_2 = \frac{1\ atm \times 1.6\ dm^3}{3\ atm}$$

$$V_2 = 0.53\ dm^3$$

The new volume of balloon is $0.55dm^3$. It will contract.

Q.8 A sample of neon gas occupies 75.0 cm^3 at very low pressure of 0.4 atm. Assuming temperature remain constant what would be the volume at 1.0 atm pressure?

Given Data:

$P_1 = 0.4$ atm
$V_1 = 75.0\ cm^3$
$P_2 = 1$ atm

Required:

$V_2 = ?$

Solution

By using the equation of Boyle's law

$$P_1V_1 = P_2V_2$$

By putting the values

$$0.4\ atm \times 75\ cm^3 = 1\ atm \times V_2$$

$$V_2 = \frac{0.4\ atm \times 75\ cm^3}{1\ atm}$$

$$\mathbf{V_2 = 30\ cm^3}$$

Q.9 A gas occupies a volume of 35.0 dm^3 at 17°C. If the gas temperature rises to 34°C at constant pressure, would you expect the volume to double? If not calculate the new volume.

Given Data:

$T_1 = 17\ ^\circ C$
$= 273 + 17 = 290\ K$
$V_1 = 35\ dm^3$
$T_2 = 34\ ^\circ C$
$= 273 + 34 = 307\ K$

Required:

$V_2 = ?$

Solution:

Volume will not be doubled because the absolute temperature is not doubled.

By using the equation of Charle's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

By putting the values

$$\frac{35\ dm^3}{290\ K} = \frac{V_2}{307\ K} \text{ or}$$

$$V_2 = \frac{35\ dm^3 \times 307\ K}{290\ K}$$

$$37\ dm^3 = V_2$$

Q.9 The largest moon of Saturn, is Titan. It has atmospheric pressure of $1.6x10^5$ Pa. What is the atmospheric pressure in atm? Is it higher than earth's atmospheric pressure?

Solution:

We know that

$$1atm = 101325\ Pa$$

Atmospheric pressure of titan in Pascal $= 1.6\times10^5$Pa.

Atmospheric pressure of titan in atm $= \dfrac{1.6\times10^5}{101325}$

$= 1.58$atm

Thus the atmosphere pressure of titan (1.58 atm) is greater than the atmospheric pressure of earth (1.0atm).

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 5: Physical States of Matter
Long Questions

Matter: *Anything that has mass and occupies space is called matter.*

Examples: Air, water, table, book etc.

Physical states of matter:

Matter exists in three physical states i.e. gas, liquid and solid. The simplest form of matter is the gaseous state. Liquids are less common and most of the matter exists as solid. These states are classified by means of two properties

(i) Shape **(ii)** Volume

5.1 GASEOUS STATE

Q.1 Write down the general properties of Gaseous state.

Ans: Gases:

The state of matter that has indefinite shape and indefinite volume is called gas.

Examples:

Hydrogen (H_2), oxygen (O_2), carbon dioxide (CO_2) etc

Typical properties of gases:

i. Diffusion

The spontaneous mixing up of molecules by random motion and collisions to form a homogeneous mixture is called diffusion.

Examples:

a. Spreading of fragrance of flower

b. Spreading of fragrance of perfume

Dependence:

Rate of diffusion depends upon the molecular mass of the gases. Lighter gases diffuse rapidly than heavier ones. e.g. H_2 diffuses four times faster than O_2 gas.

ii. Effusion:

It is escaping of gas molecules through a tiny hole into a space with lesser pressure.

Example:

When a tyre gets punctured, air effuses out.

Dependence:

Effusion depends upon molecular masses of the gases. Lighter gases effuse faster than heavier gases.

iii. Pressure

Gas molecules are always in continuous state of motion. Hence when molecules strike with the walls of the container or any other surface, they exert pressure.

The force (F) exerted per unit surface area (A) is called pressure.

$$P = \frac{F}{A}$$

SI unit of Pressure:

The SI unit of force is Newton and that of area is m^2. Hence pressure has SI unit of Nm^{-2}. It is also called Pascal (Pa)

One Pascal (Pa) = 1 Nm^{-2}

Pressure Measuring Devices:

i. Barometer is used to measure atmospheric pressure

ii. Manometer is used to measure pressure in the laboratory.

Standard Atmospheric Pressure:

It is the pressure exerted by the atmosphere at the sea level.

Definition:

It is defined as the pressure exerted by a mercury column of 760 mm height at sea level.
It is sufficient pressure to support a column of mercury 760m in height of sea level.

Different units of Pressure:

1 atm =760 mm of Hg = 760 torr (1 mm of Hg = one torr)
101325 Nm^{-2} = 101325 Pa

iv. Compressibility:

Gases are highly compressible due to empty spaces between their molecules. When the gases are compressed, the molecules come closer to one another and occupy less volume as compared to the volume in uncompressed state.

v. Mobility:

The ease of flow of liquids is called mobility.

a. Gas molecules are always in state of continuous motion.

b. They can move from one place to another because gas molecules possess very high kinetic energy.

c. They move through empty spaces that are available for the molecules to move freely.

d. The mobility or random motion results in mixing" up of gas molecules to produce a homogeneous mixture.

vi. Density of Gases:

The mass per unit volume of a substance is called density.

Units of Measurement:

Gases have low density than liquids and solids. It is due to light mass and more volume occupied by the gas molecules. Gas density is expressed in grams per dm^3, whereas, liquid and solid densities are expressed in grams per cm^3 i.e. liquids and solids are 1000 times denser than gases.

Effect of Temperature:

The density of gases increases by cooling because their volume decreases.

Example:

At normal atmospheric pressure, the density of oxygen gas is $1.4gdm^{-3}$ at 20°C and $1.5gdm^{-3}$ at 0°C.

5.2 LAWS RELATED TO GASES

Q.2 State Boyle's Law. Give the experimental verification of Boyle's law.

Ans: Boyle's Law

Introduction:

In 1662 Robert Boyle studied the relationship between the volume and pressure of a gas at constant temperature. Robert Boyle (1627-1691) was natural philosopher, chemist, physicist and inventor. He is famous for 'Boyle's law of gases.

Statement 1:

The volume of a given mass of a gas is inversely proportional to its pressure provided the temperature remains constant.

Mathematical Expression:

According to this law the volume (V) of a given mass of a gas decreases with the increase of pressure (P) and vice versa.

It can be written as:

$$V \times \frac{1}{\text{Pressure}} \qquad V = \frac{K}{P}$$

OR

$$PV = k = \text{constant}$$

Where, k is proportionality constant. The value of k is same for the same amount of a given mass.

Another form of Boyle's Law:

Boyle's Law can also be stated as:

The product of pressure and volume of a fixed mass of a gas is constant at a constant temperature.

	P_1V_1	= K, then	P_2V_2	= K
Where	P_1	=Initial pressure	P_2	= Final pressure
	V_1	= Initial volume	V_2	= Final volume

As both equations have same constant, therefore their variables are also equal to each other P_1V_2-P_2V_2. This equation establishes the relationship between pressure and volume of the gas.

Experimental Verification of Boyle's law:

The relationship between volume and pressure can be verified experimentally by the following series of experiments. Let us take some mass of a gas in a cylinder having a movable piston and observe the effect of increase of pressure on its volume. The phenomenon is represented when the pressure of 2 atmosphere (atm) is applied, the volume of the gas reads as 1 dm^3 when pressure is increased equivalent to 4 atm, the volume of the gas reduces to 0.5 dm^3, again when pressure is increased three times i.e. 6 atm, the volume reduces to 0.33 dm^3, Similarly, when pressure is increased up to 8 atm on the piston, volume of the gas decreases to 0.25 dm^3 where 'k' is proportionality constant. The value of k is same for the same amount of a given gas.

2 atm 4 atm 6 atm 8 atm
1 dm³ 0.5 dm³ 0.33 dm³ 0.25 dm³

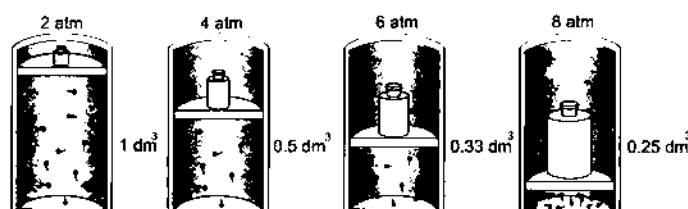



Fig. 5.1 The decrease of volume with increase of pressure.

Result

When we calculate the product of volume and pressure for this experiment, the product of all these experiments is constant i.e 2 atm dm^3. It proves the Boyle's law

$$P_1V_1 = 2atm \times 1dm^3 = 2atmdm^3$$
$$P_2V_2 = 4atm \times 0.5dm^3 = 2atmdm^3$$
$$P_3V_3 = 6atm \times 0.33dm^3 = 2atmdm^3$$
$$P_4V_4 = 8atm \times 0.25dm^3 = 2atmdm^3$$

Q.3 State Charle's Law. Give the experimental verification of Boyle's law.

Ans: Charle's Law:

Introduction

The relationship between volume and temperature keeping the pressure constant was also studied by French scientist J. Charles in 1787. J. Charles (1746-1823) was a French inventor scientist, mathematician and balloonist. He described in 1802 how gases tend to expand when heated.

Statement-1:

"The volume of a given mass of a gas is directly proportional to the absolute temperature if the pressure is kept constant"

Mathematical Expression:

When pressure P is constant, the volume V of a given mass of a gas is proportional to absolute temperature T.

Mathematically

It is represented as

$$V \times T$$

$$V = k\,T \quad \text{Or} \quad \frac{V}{T} = k$$

Where k is proportionality constant

Another form of Charle's Laws:

If temperature of the gas is increased its volume also increases. When temperature is changed from T_1 to T_2, the volume will change from V_1 to V_2. The mathematical form of Charles' Law will be:

$$\frac{V_1}{T_1} = k, \quad \frac{V_2}{T_2} = k$$

As both equations have same value of constant, therefore their variables are also equal to each other.

So

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Experimental Verification of Charles' Law:

Let us take a certain amount of gas enclosed in a cylinder having a movable piston. If the initial volume of the gas V_1 is 50 cm^3 and initial temperature T_1 is 25°C on heating the cylinder up to 100 °C its new volume V_2 is about 62.5 cm^3. The increase in temperature increases the volume that can be observed as elaborated.

Representation of increase of volume with the increase of temperature:

According to Charle's Law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{50}{25+273} = \frac{62.5}{100+273}$$

$$\frac{50}{298} = \frac{62.5}{373}$$

$$0.167 = 0.167$$

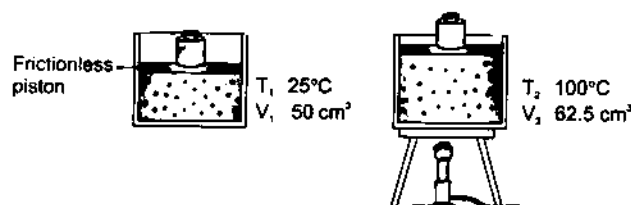



Fig. 5.2: Representation of increase of volume with the increase in temperature.

Q.4 Explain the absolute temperature scale with example.

Ans: Absolute Temperature Scale:

Introduction:

Lord Kelvin introduced absolute temperature scale or Kelvin scale. This scale of temperature starts from 0 K or $-273.15°C$, which is given the name of absolute zero.

Absolute Zero:

It is the temperature at which an ideal gas would have zero volume.

Absolute temperature scale or Kelvin scale:

A scale of temperature that starts from zero Kelvin or –273.25°C is called absolute temperature scale or Kelvin scale.

As both scales have equal degree range, therefore, when 0K is equal to -273°C then 273 K is equal to 0°C as shown in the scales

Conversion of Kelvin temperature to Celsius temperature and vice versa can be carried out as follows:

$$T\,(K) = T\,(°C) + 273$$

$$T\,(°C) = T\,(K) - 273$$

Remember

Always convert temperature scale from °C to K scale while solving problems.

K = 273 + °C

Q.5 In which units body temperatures is measured?

Ans: Body temperature is measured in Fahrenheit scale. Normal body temperature is 98.6F, it is equivalent to 37°C.

- **i.** This temperature is close to average normal atmospheric temperature.
- **ii.** In winter atmospheric temperature falls lower than that of our body temperature.
- **iii.** According to principle of heat flow, heat flows out from our body and we feel cold.
- **iv.** To control this outward flow of heat, we wear black and warm clothes.
- **v.** To maintain body temperature we use dry fruits, tea, coffee and meats etc.

Q.6 Explain the Physical States of Matter and Role of Intermolecular Forces.

Ans: Physical States of Matter and Role of Intermolecular Forces:

Matter exists in three physical states; gas, liquid and solid.

Gaseous state of matter:

In the gaseous state, the molecules are far apart from each other. Therefore, intermolecular forces are very weak in them.

Liquid state of matter:

In the liquid state molecules are much closer to each other as compared to gases. As a result liquid molecules develop stronger intermolecular forces, which affect their physical properties like diffusion, evaporation, vapour pressure and boiling point.

Compounds having stronger intermolecular forces have higher boiling points.

Solid State of Matter

The intermolecular forces become so dominant in solid state that the molecules look motionless. They arrange in a regular pattern therefore they are denser than molecules of liquids.

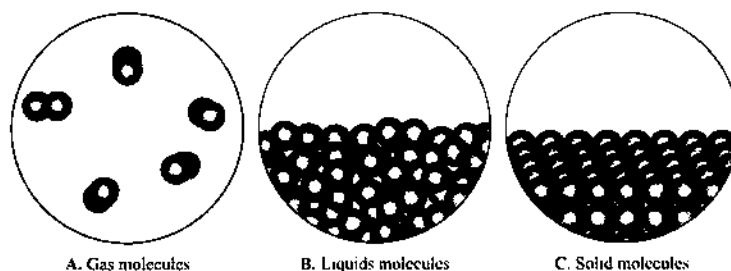



Fig 5.3 Three states of matter showing intermolecular forces.

Q.7 What are liquids? Name some important properties of liquids.

Ans: Liquid State

The state of matter that has indefinite shape but definite volume is called liquid.

Typical Properties of liquids:

- **i.** Evoporation
- **ii.** Vapour pressure
- **iii.** Boiling point
- **iv.** Freezing point
- **v.** Diffusion
- **vi.** Density

Liquids have a definite volume but their shape is not definite. A liquid attains shape of the container in which it is put. A few typical properties of the liquids are given below.

Q.8 Explain the Evaporation. Write down the factor effecting evaporation.

Ans: Evaporation:

The process of changing of a liquid into a gas phase is called evaporation.

The molecules having more than average kinetic energy over come the attractive forces among the molecules and escape from the surface is called as evaporation.

It is reverse to condensation in which a gas changes into liquid. Evaporation is an endothermic process. Evaporation is an endothermic process (heat is absorbed). Such as when one mole of water in liquid state is converted into vapour form, it requires 40.7 kJ of energy.

$$H_2O_{(l)} \longrightarrow H_2O_{(g)} \qquad \Delta H^{\circ}_{vap} = 40.7 \text{ Kj/mol}$$

Mechanism of evaporation:

In the liquid state, molecules are in a continuous state of motion. They possess kinetic energy but all the molecules do not have same kinetic energy. Majority of the molecules have average kinetic energy and a few have more than average kinetic energy. The molecules having more than average kinetic energy, overcome the attractive forces among the molecules and escape from the surface. It is called as evaporation.

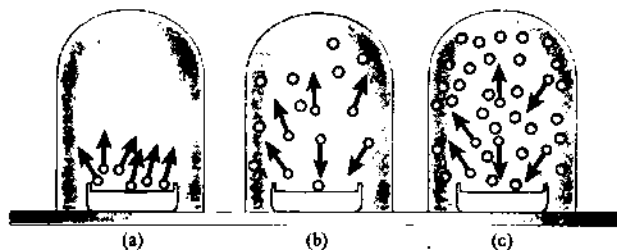



Fig. 5.4 A state of Dynamic Equilibrium between liquid and its vapours

Evaporation and Temperature:

Evaporation is a continuous process taking place at all temperatures. The rate of evaporation is directly proportional to temperature. It increases with the increase in temperature because of increase in kinetic energy of the molecules.

Evaporation is a cooling process:

When the high kinetic energy molecules vaporize, the temperature of remaining molecules

falls down. To compensate this deficiency of energy, the molecules of liquid absorb energy from the surroundings. As a result the temperature of surroundings decreases and we feel cooling.

Example:

When we put a drop of alcohol on palm, the alcohol evaporates and we feel cooling ffect.

Factors affecting evaporations:

Evaporation depends upon following factors:

i. Surface area:

Evaporation is a surface phenomenon. Greater is surface area, greater is evaporation and vice verse. Example, sometimes a saucer is used if tea is to be cooled quickly. This is because evaporation from the larger surface area of saucer is more than that from the smaller surface area of a tea cup.

ii. Temperature:

At high temperature, rate of evaporation is high because at high temperature kinetic energy of the molecules increases so high that they over- come the intermolecular forces and evaporate rapidly.

For example:

Hot water will evaporate faster than the cold water in containers of same capacity.

iii. Intermolecular forces:

The stronger the inter molecular attractive forces, the lower is the evaporation.

For Example, water has stronger intermolecular forces than alcohol, therefore, alcohol evaporates faster than water.

Temp	Vapour Pressure	Temp	Vapour Pressure
°C	mmHg	°C	mmHg
0	4.58	60	149.4
20	17.5	80	355.1
40	55.3	100	760.0

Q.9 What is vapour pressure and how it is affected by inter molecular forces? Ex. Q. 3

Ans: Vapour Pressure:

The pressure exerted by the vapours of a liquid at equilibrium with the liquid at a particular temperature is called vapour pressure of a liquid.

State of equilibrium:

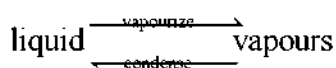
The equilibrium is a state when rate of vaporization and rate of condensation is equal to each other but in opposite directions.

Dynamic Equilibrium:

The state at which two processes take places in the opposite direction at equal rates simultaneously is called dynamic equilibrium.

The number of molecules evaporating will be equal to the number of molecules coming back (condensing) to liquid. This state is called dynamic equilibrium.

Explanation:

$$\text{liquid} \underset{\text{condense}}{\overset{\text{vapourize}}{\rightleftharpoons}} \text{vapours}$$

From the open surface of a liquid, molecules evaporate and mix up with the air but when we close a system, evaporated molecules start gathering over the liquid surface. Initially the vapours condense slowly to return to liquid. After sometime condensation process increases and a stage reaches when the rate of evaporation become equal to rate of condensation. This is called Factors affect in vapour pressure.

A state of Dynamic Equilibrium between liquid and its vapors Vapour pressure of a liquid depends upon the following factors.

. 5.6 Diffusion in liquids

i. Nature of liquid:

Vapors pressure depends upon the nature of liquid.

Polar liquids have low vapour pressure than non-polar liquids at. the same temperature.

This is because of strong intermolecular forces between the polar molecules of liquids.

Example:

Water has less vapour pressure than that of alcohol at same temperature.

ii. Size of molecules

Small size molecules can easily evaporate than big size molecules. Hence small size molecules exert more pressure.

Examples:

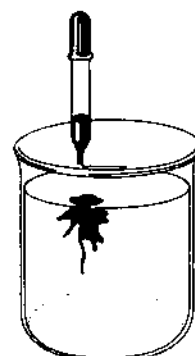
Hexane (C_6H_{14}) is a small sized molecule as compared to decane ($C_{10}H_{22}$).

C_6H_{14} evaporates rapidly and exerts more pressure than $C_{10}H_{22}$).

iii. Temperature

At high temperature, vapour pressure is higher than at low temperature. At elevated temperature, the kinetic energy of the molecules increases enough to enable them to vaporize and exert pressure.

Example:

Vapour pressure of water at 0°C is 4.58 mmHg and at 100°C it is 760 mmHg.

Q.10 Define boiling point and also explain how it is affected by different functions?

Ans: Boiling Point:

The temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure is called boiling point.

Mechanism of boiling:

When a liquid is heated, its molecules gain energy and the number of molecules which have more than average kinetic energy increases. More and more molecules become energetic, enough to overcome the intermolecular forces. Due to this, rate of evaporation increases which results in increase of vapour pressure until a stage reaches where the vapour pressure of a liquid becomes equal to atmospheric pressure. At this stage the liquid starts boiling.

Relationship between boiling point and vapour pressure:

Let's describe the increase of vapour pressure of diethyl ether, ethyl alcohol and water with the increase of temperature. At O°C the vapour pressure of diethyl ether is 200

mm Hg, of ethyl alcohol 25 mm Hg while that of water is about 5 mm Hg. When they are heated, vapour pressure of diethyl ether increases rapidly and becomes equal to atmospheric pressure at 34.6°C, while vapour pressure of water increases slowly because intermolecular forces of water are stronger.

The figure shows the vapour pressure increases very rapidly when the liquids are near to boiling point.

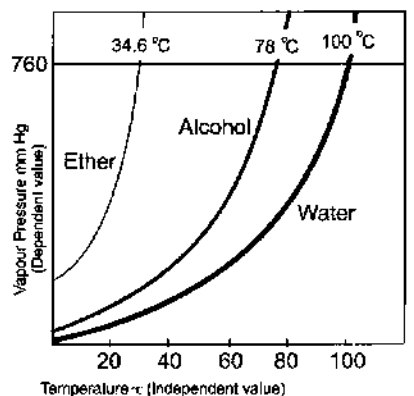



Fig.5.5 Boiling point curves of Ether Alcohol and water.

Factors Affecting Boiling Point:

The boiling point of the liquid depends upon the following factors.

i. Nature of liquid:

The polar liquids have high boiling points than that of non-polar liquids because polar liquids have strong intermolecular forces.

Example:

Boiling point of water (polar) is 100°C while that of ethyl alcohol (non polar) is 78°C.

ii. Intermolecular forces:

The stronger the inter molecular forces, the higher is the boiling point of liquid.

Intermolecular forces play a very important role on the boiling point of liquids.

Substances having stronger intermolecular forces have high boiling points, because such liquids attain a level of vapour pressure equal to external pressure at high temperature.

Example:

Boiling point of water (100°C) is greater than that of alcohol (78°C) due to stronger inter molecular forces of attraction.

iii. External pressure:

Boiling point of a liquid depends upon external pressure. Boiling point of a liquid is controlled by external pressure in such a way, that it can be increased by increasing external pressure and vice versa. This principle is used in the working of 'Pressure Cooker'.

Q.11 What is meant by freezing point?

Ans: Freezing Point:

The temperature at which vapour pressure of a liquid state becomes equal to the vapour pressure of the solid state and liquid and solid coexist in dynamic equilibrium is called freezing point.

Explanation:

When liquids are cooled the vapour pressure of liquid decreases and when vapour pressure of a liquid state becomes equal to the vapour pressure of the solid state. At this temperature liquid and solid coexist in dynamic equilibrium with one another and this is called the freezing point of a liquid.

Examples: Freezing point of water is 0°C and that of acetic acid is 16.6°C. Due to attractive force respectively.

Q.12 Describe the phenomenon of diffusion in liquids along with factors which influence it.

Ans: Diffusion:

A spontaneous mixing up of molecules by random motion and collisions to form homogenous mixture is called diffusion.

Ultimate goal:

The liquid molecules are always in a state of continuous motion. They move from higher concentration to lower concentration. They mix up with the molecules of other liquids, so that they form a homogeneous mixture.

Example:

When a few drops of ink are added in a beaker of water, ink molecules move around and after a while spread in whole of the beaker. Thus diffusion has taken place.

Comparison of rate of diffusion of liquids and gases:

Liquids diffuse like gases but the rate of diffusion of liquid is very slow.

Factors affecting diffusion:

The diffusion of liquids depends upon the following factors.

i. Intermolecular force:

Liquids having weak intermolecular forces diffuse faster than those of solids having strong intermolecular forces.

Example:

Rate of diffusion of alcohol is greater than that of water.

ii. Size of molecules:

Big size molecules diffuse slowly. For example, honey diffuses slowly in water than that of alcohol in water.

iii. Shapes of molecules:

Regular shaped molecules diffuse faster than irregular shaped molecules because they can easily slip over and move faster.

iv. Temperature:

Diffusion increases by increasing temperature because at high temperature the intermolecular forces are weak.

Example:

Rate of diffusion of water is higher at 25°C than that of 0°C.

Q.13 Explain comparison between densities of gases and liquids.

Ans: Density

"The mass per unit volume of a substance is called density."

Dependence of Density of liquids:

The density of liquid depends upon its mass and volume.

Comparison between densities of gases and liquids:

Liquids are denser than gases because molecules of liquid are closely packed and the spaces between their molecules are negligible. As the liquid molecules have strong intermolecular forces hence they cannot expand freely and have a fixed volume. Like gases, they cannot occupy all the available volume of the container that is the reason why densities of liquids are high.

Examples:

Density of water is 1.0 g cm^{-3} while that of air is 0.001 g cm^{-3} that is the reason why drops of rain fall downward.

Variation in densities of liquids:

The densities of liquids also vary. You can observe kerosene oil floats over water while honey settles down in the water.

Q.14 Explain typical properties of solid state.

Ans: Solid State:

The state of matter which has definite shape and definite volume is called solid.

Examples:

Sugar, common salt, iron, gold etc

In solid state the molecules are very close to one another and they are closely packed. The intermolecular forces are so strong that particles become almost motionless. Hence they cannot diffuse. Solid particles possess only vibrational motion.

Typical Properties of Solids:

Some typical properties of solids are as follows:

i. Melting Point:

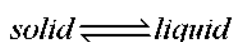
The temperature at which the solid starts melting and coexists in dynamic equilibrium with liquid state is called melting point.

Examples:

Melting point of sodium chloride is 801°C.

Explanation:

The solid particles possess only vibrational kinetic energy. When solids are heated, their vibrational energies increase and particles vibrate at their mean position with a higher speed. If the heat is supplied continuously, a stage reaches at which the particles leave their fixed positions and then become mobile. At this temperature solid melts.

$$solid \rightleftharpoons liquid$$

Melting points of Ionic and Covalent solids:

The ionic and covalent solids make network structure to form macromolecules so all such solids have very high melting points.

ii. Rigidity:

The particles of solids are not mobile. They have fixed position. Therefore solids are rigid in their structure.

iii. Density

Mass per unit volume of a substance is called density.

Comparison between densities of solids, liquids and gases:

Solids are denser than liquids and gases because solid particles are closely packed and do not have empty spaces between their particles. Therefore, they have the highest densities among the three states of matter.

Examples:

Density of aluminum is 2.70 g cm^{-3}, iron is 7.86 g cm^{-3} and gold is 19.3 g cm^{-3}.

Q.15 Differentiate between crystalline and amorphous solids

OR

Explain the types of solids in detail?

Ans: Types of Solids: (Greek word amorphous means without shape or shapeless)

According to their general appearance solids can be classified into two types:

i. Amorphous solids **ii.** Crystalline solids

i. Amorphous Solids:

"Solids in which the particles are not arranged in a regular repeating pattern are called amorphous solids".

They do not have sharp melting points. They do not form crystals.

Examples: Plastic rubber, glass, coal tar etc.

ii. Crystalline Solids:

Solids in which particles are arranged in a definite three-dimensional pattern are called crystalline solids.

Properties:

i. They have definite surfaces or faces.

ii. Each face has definite angle with the other.

iii. They have sharp melting points.

Examples: Diamond, sodium chloride, sugar, ammonium chloride etc

Q.16 Define Allotropy. Explain its condition and properties.

Ans: Allotropy:

The existence of an element in more than one form, in same physical state is called allotropy.

OR

Different forms of same element having same chemical properties but different physical properties are called allotropes or allotropic forms and this phenomenon is called allotropy.

Reasons:

i. Different number of atoms OR in a molecule:

Different forms of an element which have same chemical properties but different

physical properties are called allotropes or allotropic forms and this phenomenon is called allotropy.

The existence of two or more kinds of molecules of an element each having different number of atoms such as allotropes of oxygen are oxygen (O_2) and ozone (O_3).

ii. Different arrangement of atoms in a molecule:

Different arrangement of two or more atoms or molecules in a crystal of the element

Examples:

a. Sulphur shows allotropy due to different arrangement of molecules (S_8) in the crystals.

b. Due to different arrangement of carbon (c) atoms in the crystals carbon has three allotropes.

Diamond, Graphite, Bucky balls

c. Due to different arrangement of P_4 molecules in the crystals, phosphorous exists in the three allotropes i.e. White, Red, Black

Properties of allotropes:

They always show different physical properties but may have same or different chemical properties.

Effect of temperature:

Allotropes of solids have different arrangement of atoms in space at a given temperature.

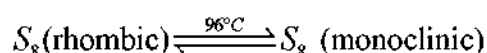
The arrangement of atoms also changes with the change of temperature and new allotropic form is produced.

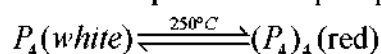
Transition temperature:

The temperature at which one allotrope changes into another is called transition temperature.

Examples:

i. Transition temperature of sulphur is 96°C, below this temperature rhombic form is stable. If rhombic form is heated up to 96 °C, its molecules rearrange themselves to give monoclinic form.

$$S_8(\text{rhombic}) \xrightleftharpoons{96°C} S_8\ (\text{monoclinic})$$

ii. Transition temperature for allotropic forms of phosphorous is 250°C.

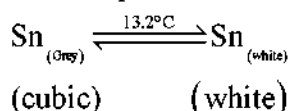
$$P_4(white) \xrightleftharpoons{250°C} (P_4)_4\,(\text{red})$$

White Phosphorous:

Is a very reactive, poisonous and waxy, soft solids. It exists as tetra-atomic molecules.

Red Phosphorous:

Is less reactive, non poisonous and brittle powder

iii. Transition temperature for allotropic forms of tin is 13.2°C.

$$\underset{(\text{cubic})}{Sn_{(\text{Grey})}} \xrightleftharpoons{13.2°C} \underset{(\text{white})}{Sn_{(\text{white})}}$$

Q.17 Explain change of instrumentation as the science progresses.

Ans: Change of instrumentation as the science progresses

There are many aspects to be considered about the functioning of instruments. Scientific observation is determined by the human sensory system. It generally relies on instruments that serve as mediators between the world and the senses. Thus, instruments can be considered as a reinforcement of the senses. They provide a great capacity for increasing the power of observation and making induction processes easier. Furthermore, scientific instruments constitute a major factor in checking, refuting or changing previously established theories.

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 5: Physical States of Matter
Extra MCQ's

1. **Which of the following expand or compressed easily**
   (a) Liquid (b) Gases (c) Solids (d) Water
2. **Diffusion is faster in:**
   (a) Liquid (b) Solids (c) Gases (d) Plasma
3. **Gases exerts pressure in all directions**
   (a) Uniformly (b) Randomly (c) Alternately (d) Constantly
4. **Which principle is used in working pressure cooker?**
   (a) Boiling point increased by increasing the external pressure
   (b) Boiling point decreased by decreasing the external pressure
   (c) Boiling point increased by increasing the external pressure
   (d) None of these
5. **Red phosphorus is:**
   (a) Less reactive (b) Non-poisonous (c) Brittle (d) All of above
6. **Which percentage of salt is required to kill the bacteria?**
   (a) 10% (b) 20% (c) 30% (d) 40%
7. **The density of gold is:**
   (a) 2.70 gcm^{-3} (b) 7.86 gcm^{-3} (c) 9.3 gcm^{-3} (d) 2.98 $gcmw^{-3}$
8. **The solids in which particles are arranged in definite three-dimensional pattern are called:**
   (a) Solids (b) Crystalline solids(c) Amorphous solids(d) both 'B' and 'C'
9. **The transition temperature of tin is:**
   (a) 96°C (b) 250°C (c) 18°C (d) 100°C
10. **Oxygen has two allotropic forms:**
    (a) O_2 and O_4 (b) O_2 and O_3 (c) O and O_3 (d) O_2 and O
11. **Plastics, glass rubber, lamp-black etc. are the examples of:**
    (a) Crystalline solids (b) Super cooled liquids
    (c) Amorphous solids (d) Ionic solids
12. **The temperature at which the solid and liquid phases of a substance co-exist:**
    (a) Freezing point (b) Boiling point (c) Melting point (d) Equilibrium
13. **The existence of solid in different physical forms is called:**
    (a) Crystals (b) Allotropy (c) Evaporation (d) Transition
14. **It depends upon the nature of liquid, intermolecular forces and external pressure.**
    (a) Melting point (b) Freezing point (c) Boiling point (d) None of these
15. **The conversion of a liquid into vapours at all temperature is called.**
    (a) Evaporation (b) Boiling (c) Cooling process (d) Both 'A' and 'C'
16. **Density is expressed in:**
    (a) g cm^{-3} (b) g dm^{-3} (c) Both 'A' and 'B' (d) None of these
17. Charles described, how gases tend to expand when heated:
    (a) 1882 (b) 1802 (c) 1820 (d) 1828
18. **Mass per unit volume is called:**

(a) Pressure (b) Temperature (c) Density (d) Solubility

19. Robert Boyle's was a natural:
(a) Philosopher (b) Chemist (c) Physicist (d) All of above

20. Matter exists in three physical states:
(a) Solid, liquid, plasma (b) Solid, water, vapour
(c) Plasma, liquid, solid (d) Solid, liquid, gas

21. Which one of the following gases diffuse most slowly?
(a) O_2 (b) Cl_2
(c) NH_3 (d) H_2

22. Pascal is the SI unit of:
(a) Force (b) Volume
(c) Momentum (d) Pressure

23. At 0°C the density of O_2 gas is:
(a) 1.5g dm^{-3} (b) 1.4g dm^{-3}
(c) 4.1g dm^{-3} (d) 0.15g dm^{-3}

24. $\frac{V}{T} = K$ is the mathematical form of:
(a) Boyle's law (b) Charles law
(c) Avogendo law (d) Dalton law

25. –273°C after conversion to Kelvin scale becomes:
(a) OK (b) 273K
(c) 173K (d) 100K

26. At which temperature rate of evaporation of water is minimum?
(a) 50°C (b) 40°C
(c) 90°C (d) 70.5°C

27. The molecular formula of decane is
(a) C_6H_{14} (b) C_7H_{16}
(c) CaH_{20} (d) $C_{10}H_{22}$

28. C_6H_{14} is the molecular formula of
(a) Pentane (b) Hexane
(c) Propane (d) Decane

29. At a given temperature which of the following has maximum vapour pressure
(a) C_5H_{12} (b) C_6H_{14}
(c) C_7H_{10} (d) C_8H_{18}

30. The density of gold is:
(a) 270 gcm^{-3} (b) 7.86g cm^{-3}
(c) 19.3g cm^{-3} (d) 1.4g cm^{-3}

31. 250°C is the transition temperature of which substance:
(a) Tin (b) Carbon
(c) Phosphorus (d) Sulphur

32. The crystal structure of which tin is
(a) cubic (b) tetragonal
(c) monoclinic (d) None of these

33. Freezing point of acetic acid is:
(a) –116°C (b) 16.6°C
(c) –56°C (d) –115°C

34. The vapour pressure of which liquid is 200mm Hg at 0°C?
(a) Water (b) Alcohol
(c) Honey (d) Ether

35. 2×101325 Pa is equal to

(a) 1.56 atm
(b) 2.0 atm
(c) 15.8 atm
(d) 8.15 atm

36. Normal body temperature is
(a) 98.6°F
(b) 37°C
(c) Both a and b
(d) 98.6°C

37. The freezing point of n– octane is
(a) –57°C
(b) 126°C
(c) 34.6°C
(d) –116°C

38. 760 mmHg is equal to
(a) 266 torr
(b) 2660 torr
(c) 626 torr
(d) 1atm

39. 1 atm has how many pascals?
(a) 151987.5 Pa
(b) 19187.5 Pa
(c) 9187.6 Pa
(d) 101325Pa

40. In which unit density of gases is measures?
(a) kg mm^{-3}
(b) cm^{-3}
(c) gdm^{-3}
(d) gcm^{-3}

ANSWER KEY

1	b	11	c	21	b	31	c
2	c	12	a	22	d	32	d
3	a	13	b	23	a	33	b
4	c	14	c	24	b	34	b
5	d	15	a	25	a	35	b
6	b	16	c	26	b	36	c
7	c	17	b	27	d	37	a
8	b	18	c	28	b	38	d
9	c	19	d	29	a	39	d
10	b	20	d	30	c	40	c

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 6: Solutions
Exercise Questions

Exercise Multiple Choice Question Answers

1. **Mist is an example of solution:**
   (a) Liquid in gas (b) Gas in liquid (c) Solid in gas (d) Gas in solid
2. **Which one of the following is a 'liquid in solid' solution?**
   (a) Sugar in water (b) Butter (c) Opal (d) Fog
3. **Concentration is ratio of:**
   (a) Solvent to solute (b) Solute to solution (c) Solvent to solution (d) Both 'A' and 'B'
4. **Which one of the following solutions contains more water?**
   (a) 2M (b) 1M (c) 0.5 M (d) 0.25 M
5. **A 5 percent (w/w) sugar solution means that:**
   (a) 5 g of sugar is dissolved in 90 g of water
   (b) 5 g of sugar is dissolved in 100 g of water
   (c) 5 g of sugar is dissolved in 105 g of water
   (d) 5 g of sugar is dissolved in 95 g of water
6. **If the solute-solute forces are strong enough than those of solute-solvent forces. The solute:**
   (a) Dissolves readily (b) Does not dissolve
   (c) Dissolves slowly (d) Dissolves-and precipitates.
7. **Which one of the following will show negligible effect of temperature on its solubility?**
   (a) KCl (c) $NaNO_3$ (b) KNO_3 (d) NaCl
8. **Which one of the following is heterogeneous mixture?**
   (a) Milk (b) ink (c) Milk of magnesia (d) Sugar solution
9. **Tyndall effect is shown by:**
   (a) Sugar solution (c) jelly (b) Paints (d) Chalk solution
10. **Tyndall effect is due to:**
    (a) Blockage of beam of light (b) Non-scattering of beam of light
    (c) Scattering of beam of light (d) Passing through beam of light
11. **If 10 cm^3 of alcohol is dissolved in 100 g of water, it is called:**
    (a) % w/w (b) %w/v (c) % v/w (d) %v/v
12. **When a saturated solution is diluted it turns into:**
    (a) Supersaturated solution (b) Saturated solution
    (c) A concentrated solution (d) Unsaturated solution
13. **Molarity is the number of moles of solute dissolved in:**
    (a) lkg of solution (b) 100 g of solvent (c) 1 dm^3 of solvent (d) 1 dm^3 of solution.

ANSWR KEY

1	a	3	b	5	b	7	d	9	c	11	c	13	d

2	b	4	d	6	c	8	c	10	c	12	d	

Exercise Short Question Answers

Q.1 Why suspensions and solutions do not show Tyndall effect, while colloids do?

Ans: Suspensions and solutions do not show Tyndall effect because in suspensions particles are so big that light is blocked and difficult to pass. But in solution particles are so small that they cannot scatter the rays of light, thus do not show Tyndall effect. But colloids can show Tyndall effect because particles scatter the path of light rays thus emitting the beam of light i.e., exhibit the Tyndall effect.

Q.2 What is the reason for the difference between solutions, colloids and suspensions?

Ans: The differentiation between solutions, colloids and suspensions is based upon the particle size. In colloidal solutions the particles size is intermediate between true solutions and suspensions.

Q.3 Why does not the suspension form a homogeneous mixture?

Ans: In suspension particles remain un-dissolved due to their big size. After sometime particles settle down under the action of gravity, therefore suspension does not forma homogeneous mixture.

Q.4 How will you test whether given solution is a colloidal solution or not?

Ans: We will pass light in the solution, if the given solution scattered the light then it is a colloidal solution. It solution does not scatter the light then it is not colloidal solution.

Q.5 Classify the following into true solution and colloidal solution Blood, starch solution, glucose solution, tooth paste, copper sulphate solution, silver nitrate solution.

Ans:

True Solutions	Colloidal Solution
Glucose solution, Copper sulphate solution, silver nitrate solution	Blood, tooth paste, starch solution

Q.6 Why we stir paints thoroughly before using?

Ans: Paints are heterogeneous mixture of un-dissolved particles in a given medium. Particles settle down after sometime. So we stir paints to mix thoroughly before using.

Q.7 Which of the following will scatter light and why? Sugar solution, soap solution and milk of magnesia.

Ans: **Soap solution:**

Soap solution will scatter light (Tyndall effect) because it is colloidal solution and its particles are large enough to scatter the light.

Sugar Solution:
Sugar solution will not scatter light because the particles of sugar solution are so small that they cannot scatter light.
Milk of Magnesia:
Milk of magnesia cannot scatter the light because it is suspension and its particles are so big that light is blocked.

Q.8 What do you mean by "like dissolves like?" Explain with examples.
Ans: "Like dissolves like" means that polar substances are dissolved in polar solvents and non polar substances are soluble in non polar solvents.
For example: NaCl (polar) dissolves in water (polar solvent) and does not dissolve in benzene (non polar)

Q.9 How does nature of attractive forces of solute-solute and solvent-solvent affect the solubility?
Ans: Solubility depends upon solute solvent attractions. If the attractive forces between solvent are stronger the solubility is greater. If the attractive forces become weaker in solute there will be greater solubility.
If the attractive forces between solute particles are stronger than solute solvent forces, solute remains insoluble and solution is not formed.

Q.10 How you can explain the solute-solvent interaction to prepare a NaCI solution?
Ans: When NaCl is added in water it dissolves readily because the attractive forces between the ions of NaCl and polar molecules of water are strong enough to overcome the attractive forces between Na^+ and Cl^- ions in solid NaCl crystal. In this process, positive end of the water dipole is oriented towards the Cl^- ions and the negative end of water dipole is oriented towards the Na^+ ions. These ion-dipole attractions between Na^+ ions and water molecules, Cl^- ions and water molecules are so strong that they pull these ions from their positions in the crystal and thus NaCl dissolves.

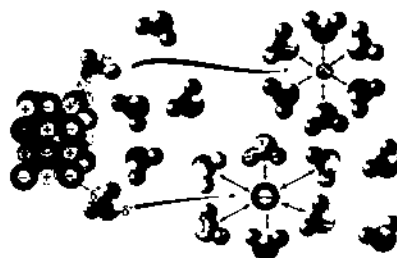



Fig. 6.2 Inter-action of solute and solvent to form solution.

Q.11 Justify with an example that solubility of a salt increases with the increase in temperature
Ans: Solubility of some salts which are usually ionic in nature increases with the increase in temperature for such solutes. It means that heat is required to break the attractive forces between the ions of solute. This process is called endothermic.
For example:
Solubility of KNO_3 and KCl can be enhanced by increasing temperature.

Q.12 What do you mean by volume/volume %?

Ans: It is the volume in cm^3 of a solute dissolved in 100 cm^3 of the solution.

For example:

30% of alcohol solution means 30 cm^3 of alcohol dissolved in sufficient amount of water, so that the total volume of the solution becomes 100 cm^3.

$$\frac{\%\text{Volume}}{\%\text{Volume}} = \frac{\text{Volume of solute}\left(cm^3\right)}{\text{Volume of solutio}\left(cm^3\right)} \times 100$$

Exercise Long Question Answers

Q.1 What is saturated solution and how it is prepared?

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Q.2 Differentiate between dilute and concentrated solutions with a common example.

Ans: See Q. No. **4** (Subjective Part, Long Questions Answers)

Q.3 Explain, how dilute solutions are prepared from concentrated solutions?

Ans: See Q. No. **7** (Subjective Part, Long Questions Answers)

Q.4 What is molarity and give its formula to prepare molar solution?

Ans: See Q. No. **6** (Subjective Part, Long Questions Answers)

Q.5 Explain the solute-solvent interaction for the preparation of solution.

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.6 What is general principle of solubility?

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.7 Discuss the effect of temperature on solubility.

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.8 Give the five characteristics of colloid.

Ans: See Q. No. **10** (Subjective Part, Long Questions Answers)

Q.9 Give at least five characteristics of suspension

Ans: See Q. No. **10** (Subjective Part, Long Questions Answers)

Exercise Solved Numerical

Q.1 A solution contains 50 g of sugar dissolved in 450 g of water. What is concentration of this solution?

Given Data:

Mass of sugar solute = 50g

Mass of water solvent = 450g

Required:

Concentration of solution (% m/m) = ?

Solution:

$$\% \text{ m/m} = \frac{\text{Mass of solute}(g)}{\text{Mass of solute}(g) + \text{Mass of solvent}(g)} \times 100$$

Solution:

$$\% \text{ m/m} = \frac{50g}{50g + 45g} \times 100$$

$$= \frac{50g}{500g} \times 100$$

Thus,

% m/m = 10% m/m

Q.2 If 60 cm^3 of alcohol is dissolved in 940 cm^3 of water, what is concentration of this solution?

Given Data:

Volume of alcohol solute = v = 60 cm^3
Volume of water solvent = v = 940 cm^3

Required Data:

Concentration of solution (% v/v) = ?

Formula:

$$\% v/v = \frac{\text{volume of solute}(cm^3)}{\text{volume of solute}(cm^3) + \text{volume of solvent}(cm^3)} \times 100$$

Solution:

$$\% \text{ v/v} = \frac{60cm^3}{60 \text{ cm}^3 + 940cm^3} \times 100$$

$$= \frac{60 \text{ cm}^3}{1000cm^3} \times 100$$

Thus **% v/v = 6%v/v**

Q.3 How much salt will be required to prepare following solutions (atomic mass: K=39; Na=23; S=32; O=16 and H=I)

(a) 250 cm^3 of KOH solution of 0.5 M

(b) 600 cm^3 of $NaNO_3$ solution of 0.25 M

(c) 800 cm^3 of Na_2SO_4 solution of 1.0 M

Ans:

(a) 250cm^3 of KOH solution of 0.5M

Given Data:

Molarity of solution = (M) = 0.5 M

Volume of solution = 250 cm^3 = $\frac{250}{1000}$ dm^3=0.25dm^3

Molar mass of KOH = 39+16+1=56gmol^{-1}

Required Data:

Mass of KOH =?

Solution:

$$\text{Molarity}=\frac{\text{Mass of solute}(g)}{\text{Molar mass of solute }\left(gmol^{-1}\right)\times\text{volume of solution }\left(dm^{3}\right)}$$

$$0.5M=\frac{\text{Mass of solute}(g)}{56g\ mol^{-1}\times0.25dm^{3}}$$

$$\text{Mass of solute}=0.5\times56\times0.25$$

$$\mathbf{=7g}$$

(b) 600cm³ of $NaNO_3$ solution of 0.25M

Given Data:

Molarity of $NaNO_3$ solutoin = (M) = 0.25M

Volume of solution = 600 cm³ $=\frac{600}{1000}=0.6dm^3$

Molar mass of $NaNO_3$ = 23 + 14 + 3(16)

= 85gmol⁻¹

Required:

Amount of $NaNO_3$ = m =?

Solution:

Using the formula:

$$\text{Molarity}=\frac{\text{Mass of solute}(g)}{\text{Molarmass of solute}\left(gmol^{-1}\right)\times\text{Volume of solution}\left(dm^{3}\right)}$$

$$\text{Molarity}=\frac{\text{Mass of solute}(g)}{85gmol^{-1}\times0.6dm^{3}}$$

$$\text{Mass of solute}=0.25\times85\times0.6$$

$$\text{Mass of solute}=12.75g$$

(c) 800cm³ of $Na_2\ SO_4$ solution of 1.0M

Given Data:

Molarity of Na_2SO_4 solution = M = 1 M

Volume of solution = V = 800 cm³ $=\frac{800}{1000}=0.8dm^3$

Molecular mass of Na_2SO_4 = 2(23) + 32 +4(16)

= 46 + 32 + 64

= 142gmol⁻¹

Required:

Mass of Na_2SO_4 =?

Solution:

Using the formula

$$\text{Molarity}=\frac{\text{Mass of solute}(g)}{\text{Molar mass of solute}\left(gmol^{-1}\right)\times\text{Volume of solution}\left(dm^{3}\right)}$$

$$1.0M=\frac{\text{Mass of solute}}{142gmol^{-1}\times0.8dm^{3}}$$

$$\text{Mass of solute} = 1.0 \times 142 \times 0.8$$
$$= 113.6\text{g}$$

Q.4 When we dissolve 20 g of NaCl in 400 cm^3 of solution, what will be its molarity?

Given Data:

Mass of NaCl = 20g

Molar mass of NaCl = 23 + 35.5 = 58.5gmol^{-1}

Volume of Solution = 400 cm^3 = $= \frac{400}{1000} 0.4\text{dm}^3$

Required:

Molarity of solution = ?

Solution:

Using the formula:

$$\text{Molarity} = \frac{\text{Mass of solute}(\text{g})}{\text{Molar mass of solute}(\text{gmol}^{-1}) \times \text{Volume of solution}(\text{dm}^3)}$$

$$= \frac{25\text{g}}{58.5\text{mol} \times 0.4(\text{dm}^3)}$$

$$= \frac{20}{23.4} = 0.85\text{M}$$

Q.5 We desire to prepare 100 cm^3 0.4 M solution of Mg Cl_2, how much Mg Cl_2 is needed?

Given Data:

Molarity of solution = 0.4 M

Volume of Solution = 100cm^3 $= \frac{100}{1000}\text{dm}^3 = 0.1\text{dm}^3$

Mass of $MgCl_2$ = 24 + 2(35.5) = 95g

=24+71=95gmol^{-1}

Required:

Mass of $MgCl_2$ = ?

Solution:

Using the formula:

$$\text{Molarity} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}(\text{gmol}^{-1}) \times \text{Volume of solutoin}(\text{dm}^3)}$$

$$0.4\text{M} = \frac{\text{Mass of solute}(\text{g})}{95\text{g mol}^{-1} \times 0.1\text{dm}^3}$$

Mass of solute = $0.4 \times 95 \times 0.1$

= 3.8g

Q.6 12M H_2S0_4 solutions is available in the laboratory. We need only 500cm^3 of 0.1 M solution, how it will be prepared?

Given Data:

Molarity of concentrated H_2SO_4 solution = M_1 = 12 M
Molarity of dilute H_2SO_4 solution = M_2 = 0.1 M
Volume of dilute H_2SO_4 solution = V_2 = 500cm^3

Required:

Volume of concentrated H_2SO_4 solution = V_1 = ?

Solution:

i. Determination of volume of concentrated solution:

Concentrated solution =Dilute solution

$$M_1 V_1 = M_2 V_2$$

$$12 \times V_1 = 0.1 \times 500$$

$$V_1 = \frac{0.1M \times 500\ cm^3}{12M}$$

Thus,

4.16 cm^3

ii. Preparation of solution

We take 4.16cm^3 of concentrated 12M H_2SO_4 solution with the help of graduated pipette

and put in a measuring flask of 500cm^3. Add water upto the mark, present at the neck o

flask. Now it is 0.1 molar solution of H_2SO_4 .

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 6: Solutions
Long Questions

6.1 Solutions

Q:1 Explain the term solution with the help of examples.

Ans: Solution:

"A homogeneous mixture of two or more substances is called solution."

Solute + Solvent = Solution

Examples:

i. Sugar solution
ii. Sodium chloride solution
iii. Copper sulphate solution
iv. Air
v. Brass
vi. Sea water

Physical states of solutions:

i. Solid: e.g. alloy
ii. Liquid: e.g. sea water
iii. Gas: e.g. air

Properties of a solution:

i. A solution has only one phase
ii. A solution is usually named after the name of solute
iii. The physical state of solution is the same as that of solvent
iv. It shows the properties of its components
v. It has a uniform composition
vi. It is transparent though it may be coloured
vii. It is electrically neutral

Homogeneous mixture:

"A mixture having uniform composition throughout is called homogeneous mixture."

The boundaries of the components can't be distinguished i.e. a solution exist as one phase.

Example:

i. The air we breathe is a solution of several gases.
ii. Brass is a solid solution of Zn and Cu.
iii. Sugar dissolved in water.

Q:2 Define the terms.

i. Aqueous solution **ii. Universal solvent** **iii. Solute**
iv. Solvent

Ans: **i. Aqueous Solution**

"The solution which is formed by dissolving a substance in water is called an aqueous solution."

In aqueous solutions water is always present in greater amount and termed as solvent Water is called a universal solvent because it dissolves majority of compounds present in earth's crust.

Examples:

i. Sugar in water.
ii. Table salt in water.

Distinguish between solution & pure liquid:

i. The simplest way to distinguish between a solution and a pure liquid is evaporation.
ii. The liquid which evaporates completely, leaving no residue, is a pure compound,.
iii. While a liquid which leaves behind a residue on evaporation is solution.
iv. An alloy like brass or bronze is also a homogeneous mixture.
v. Although it cannot be separated by physical means.
vi. It shows the properties of its components and
vii. It has a variable composition.

ii. Universal solvent:

"Water is called a universal solvent because it dissolves majority of compounds present in earth's crust."

Water can dissolve ionic as well as covalent compounds e.g NaCl, Cl_2, HCl

iii. Solute:

"The component of solution which is present in smaller quantity is called solute".

Examples:

A solute is dissolved in a solvent to make a solution in sugar. In sugar solution, sugar is solute and in sodium chloride solution, sodium chloride is solute.

Number of solutes present in a solution:

In a solution if more than two substances are present, one substance acts as solvent and others behave as solutes

Example:

In soft drinks, water is solvent while other substances like sugar, salts and CO_2 are solutes.

iv. Solvent:

"The component of a solution which is present in larger quantity is called solvent."

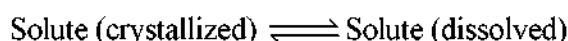
Example:

In soft drinks, water is solvent while other substances like sugar, salts and CO_2 are solutes.

Q:3 Explain saturated, unsaturated solution and super saturated solutions with the help of examples.

Ans: **i. Saturated Solution:**

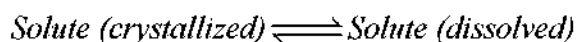
"A solution containing maximum amount of solute at a given temperature is called saturated solution"

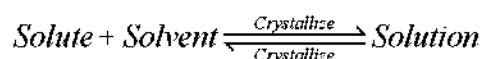
$$\text{Solute (crystallized)} \rightleftharpoons \text{Solute (dissolved)}$$

Preparation of saturated solution:

When a small amount of solute is added in a solvent, solute dissolves very easily in solvent. If the addition of solute is kept on, a stage is reached when solvent cannot dissolve more solute. At this stage, further added solute remains un-dissolved and it settles down at the bottom of the container.

Dynamic equilibrium in saturated solution:

On the particle level, a saturated solution is the one, in which un-dissolved solute is in equilibrium with dissolved solute.

$$\textit{Solute (crystallized)} \rightleftharpoons \textit{Solute (dissolved)}$$

At this stage dynamic equilibrium is established. Although dissolution and crystallization continues at a given temperature, but the net amount of dissolved solute remains constant.

$$Solute + Solvent \underset{Crystallize}{\overset{Crystallize}{\rightleftharpoons}} Solution$$

ii. Unsaturated Solution:

"A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called unsaturated solution".

Such solutions have the capacity to dissolve more solute to become a saturated solution.

Example:

$1dm^3$ of NaCl solution in which 3g of NaCl are dissolved

iii. Super saturated solution:

"The solution that is more concentrated than a saturated solution is known as supersaturated solution".

Preparation of Supersaturated Solution:

When saturated solutions are heated, they develop further capacity to dissolve more solute. Such solutions contain greater amount of solute than is required to form a saturated solution and they become more concentrated. Super- saturated solutions are not stable. Therefore, an easy way to get a supersaturated solution is to prepare a saturated solution at high temperature. It is then cooled to a temperature where excess solute crystallizes out and leaves behind a saturated solution.

Example:

A saturated solution of sodium thiosulphate ($Na_2S_2O_3$) in water at 20 °C has 20.9 g of salt per 100 cm^3 of water. Less than this amount of salt per 100 cm^3 of water at 20 ^{0}C will be an unsaturated solution. A solution having more amount than 20.9 g of salt per 100 $cm^3$of water at20 ^{0}C will be a supersaturated solution.

Q:4 Differentiate between dilute and concentrated solutions with a common example.

Ans: Dilute solutions:

Dilute solutions are those which contain relatively small amount of dissolved solute in the solution.

Example:

A solution containing 5g of sodium chloride in 100g water is a dilute solution.

Concentrated solutions:

Concentrated solutions are those which contain relatively large amount of dissolved solute in the solution.

Example

i. Brine is a concentrated solution of common salt in water. Addition of more solvent will dilute the solution and its concentration decreases.

ii. 0.1M Na_2CO_3 solution is dilute solution as compared to 5M Na_2CO_3 solution.

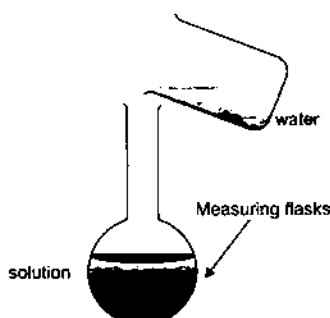



Fig. 6.1 Dilution of a solution.

Explain different types of solution with examples.

Ans: **Solution:**

A homogeneous mixture of two or more substances is called a solution

Examples: Sugar solution, Air etc

Types of Solutions:

Each solution consists of two components, solute and solvent. The solute as well as solvent may exist as gas, liquid or solid. There are nine types of solutions depending upon the physical state of solute and solvent.

Table: Different types of solutions with examples

Sr. No	Solute	Solvent	Example of Solutions
1	Gas	Gas	Air, mixture of H_2 and He in weather balloons, mixture of N_2 and O_2 in cylinders for respiration.
2	Gas	Liquid	Oxygen in water, carbon dioxide in water.
3	Gas	Solid	Hydrogen adsorbed on palladium.
4	Liquid	Gas	Mist, fog, liquid air pollutants.
5	Liquid	Liquid	Alcohol in water, benzene and toluene.
6	Liquid	Solid	Butter, cheese.
7	Solid	Gas	Dust particles or smoke in air.
8	Solid	Liquid	Sugar in water.
9	Solid	Solid	Metal alloys such as brass (Cu + Zn), bronze (Cu + Sn), opals etc

Q:5 Write down the types and properties of concentration units for solution.

Ans: Concentration Units:

Concentration:

The proportion of a solute in a solution is called concentration.

OR

It is also a ratio of amount of solute to the amount of solution or ratio of amount of solute to amount of solvent is called concentration of solution.

$$\text{Concentration of solution} = \frac{\text{Amount of solute}}{\text{Amount of solution or amount of solvent}}$$

Independence of Concentration:

Concentration does not depend upon the total volume or total amount of the solution.

Example:

A sample taken from the bulk solution will have the same concentration.

Concentration Units:

There are various types of units used to express concentration of solutions.

Percentage:

The number of parts of a component present in 100 parts of a substance is called percentage or percentage composition.

Percentage unit of concentration refers to the percentage of solute present in a solution. The percentage of solute can be expressed by mass or by volume. It can be expressed in terms of percentage composition by four different ways.

i. Percentage mass / mass (%m/m):

"It is the number of grams of solute in 100 grams of solution."

Example:

10%m/m sugar solution means that 10g of sugar is dissolved in 90g of water to make 100g of solution.

$$\%\text{m / m} = \frac{\text{Mass of solute (g)}}{\text{Mass of solution (g)}} \times 100$$

ii. Percentage - mass/volume (%m/v):

It is the number of grams of solute dissolved in 100 cm^3 of solution.

Example:

10 % m/v sugar solution contains 10 g of sugar in 100 cm^3 of solution. The exact volume of solvent is not mentioned or it is not known.

$$\%\text{m/v} = \frac{\text{Mass of solute (g)}}{\text{Volume of solution }(\text{cm}^3)} \times 100$$

iii. Percentage - volume/mass (%v/m)

It is the volume in cm^3 of a solute dissolved in 100 g of the solution.

Example:

For example, 10 % alcohol solution in water means 10 cm^3 of alcohol is dissolved in (unknown) volume of water so that the total weight of solution is 100 g. In such solutions the mass of solution is under consideration, total volume of the solution is not considered.

$$\%\text{m/v} = \frac{\text{Volume of solute }(\text{cm}^3)}{\text{Mass of solution (g)}} \times 100$$

iv. Percentage – volume/volume (% v/v)

It is the volume in cm^3 of a solute dissolved per 100 cm^3 of the solution".

Example:

30 percent v/v alcohol solution means 30 cm^3 of alcohol dissolved in sufficient amount

Formula:

$$(\%v/v) = \frac{\text{Volume of solute}(cm^3)}{\text{Volume of solution}(cm^3)} \times 100$$

Q:6 What is molarity and give its formula to prepare molar solution?

Ans: Molarity

Number of moles of solute dissolved in one dm^3 of solution is called molarity. It is represented by M.

Significance:

It is a concentration unit. The formula used for preparation of molar solution is as follows. molarity is the unit mathly used in chemistry and allied silences.

Formula:

$$\text{Molarity} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in } cm^3}$$

$$\text{Molarity (M)} = \frac{\text{No of moles of solute}}{\text{Volume of solution } (dm^3)}$$

Units of Molarity:

$$\text{Molarity} = \frac{\text{No of moles of solute}}{\text{Volume of solution in } dm^3}$$

$$M = \frac{Mol}{dm^3}$$

$$M = Moldm^{-3}$$

Preparation of Molar Solution:

A solution which contains 1mole of solute dissolved per dm^3 of solution is called molar solution.

One Molar solution is prepared by dissolving 1 mole (molar mass) of the solute in sufficient amount of water to make the total volume of the solution up to 1dm^3 in a measuring flask.

Example:

1M solution of NaOH is prepared by dissolving 40g of NaOH in sufficient water to make the total volume 1dm^3.

Relationship between solute and molarity:

Solute $\propto$ molarity

As amount of solute is increased, its concentration or molarity also increases. 2M solution is more concentrated than 1M solution.

Q:7 Explain how dilute solution are prepared from concentrated. OR
Explain dilution of solution in detail.

Ans: Dilution of Solution:

The process of decreasing concentration of solution by a adding more solvent in it is called dilution of solution.

Example:

We do have 2M solution of NaCl. If we add more solvent (water) to it, the concentration of solution decreases. This process is called dilution of solution. Dilute molar solution is prepared from a concentrated solution of known molarity.

Example:

Suppose we are to make $100cm^3$ of 0.01 M solution from given 0.1 M solution of potassium permanganate.

Given Data

$$M_1 = 0.1\,M$$
$$V_2 = 100\ cm^3$$
$$M_2 = 0.01\,M$$

Required Data:

$V_1 = ?$

Solution:

i. Determination of volume of concentrated solution:

First 0.1 M solution is prepared by dissolving 15.8 g of potassium permanganate in 1 dm^3 of solution. Then 0.01 M solution is prepared by the dilution according to following calculations:

Preparation of 0.01M, $100cm^3$ $KnmO_4$ solution =?

$$V_1 = ?$$

Putting the values in equation we get:

Concentrated solutions = dilute solution:

$$V_1 \times 0.1 = 0.01 \times 100$$
$$V_1 \times 0.1 = 0.01 \times 100$$
$$V1 = 0.01 \times 100 = 10cm^3$$

Concentrated solution of $KMnO_4$ has dense purple colour.

ii. Preparation of Solution:

We take 10 cm^3 of this solution with the help of a graduated pipette and put in a measuring flask of 100 cm^3. Add water upto the mark present at the neck of the flask. Now it is 0.01 molar solution of $KMnO_4$.

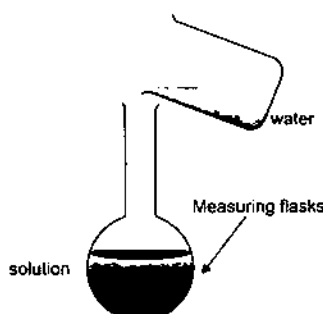



Fig. 6.1 Dilution of a solution.

Q:8 What is solubility? Write down the factors affecting solubility.

Or

What is general principle of solubility? (Ex. Q.6)

Explain solute – solvent interactions in the preparation of a solution. (Ex.Q.5)

Discuss the effect of temperature on solubility. (Ex.Q.7)

Ans: Solubility

The number of grams of the solute dissolved in 100 g of solvent to prepare a saturated solution at a particular temperature.

The concentration of a saturated solution is referred to as solubility of the solute in a given solvent.

Factor affecting the solubility

Following are the factors which affect the solubility of solutes:

a. Nature of solute and solvent (like dissolves like)

b. Solute-solvent interactions

c. Temperature

a. Like dissolves like (Nature of solute and solvent):

The general principle of solubility is, like dissolves like.

i. The polar substances are soluble in polar solvents. Ionic solids and polar covalent compounds are soluble in water

Examples: KCl, Na_2CO_3, $CuSO_4$, sugar and alcohol are soluble in water

ii. Non-polar substances are not soluble in polar solvents.

Examples: Ether, benzene and petrol are all in soluble in water

iii. Non-polar covalent substances are soluble in non-polar solvents (mostly organic solvents).

Examples: Grease, paints, naphthalene are soluble in ether or carbon tetrachloride etc.

b. Solubility and Solute-solvent interactions:

The solute-solvent interaction can be explained in terms of creation of attractive forces between the particles of solute and those of solvent.

Steps to dissolve solute in solvent:

To dissolve one substance (solute) in another substance (solvent) following three events must occur:

i. Solute particles must separate from each other

ii. Solvent particles must separate to provide space for solute particles.

iii. Solute and solvent particles must attract and mix up.

Dependence of solution formation:

Solution formation depends upon the relative strength of attractive forces between solute- solute, solvent-solvent and solute-solvent.

Physical states of solute and nature of interactions:

i. Generally solutes are solids. Ionic solids are arranged in such a regular pattern that the inter-ionic forces are at a maximum.

ii. If the new forces between solute and solvent particles overcome the solute-solute attractive forces, then solute dissolves and makes a solution.

iii. If forces between solute particles are strong enough than solute-solvent forces, solute remains insoluble and solution is not formed.

iv. The solvent molecules first pull apart the solute ions and then surround them. In this

way solute dissolves and solution forms.

Example: (Dissolution of sodium chloride)

When NaCl is added in water it dissolves readily because the attractive interaction between the ions of NaCl and polar molecules of water are strong enough to overcome the attractive forces between Na^+ and Cl^- ions in solid NaCl crystal. In this process the positive end of the water dipole is oriented towards the Cl^- ions and the negative end of water dipole is oriented towards the Na^+ ions. These ion-dipole attractions between Na^+ ions and water molecules, Cl^- ions and water molecules are so strong that they pull these ions from their positions in the crystal and thus NaCl dissolves.

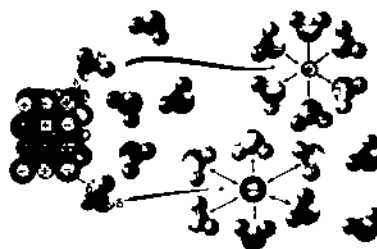



Fig. 6.2 Inter-action of solute and solvent to form solution.

Q:9 Discuss the effect of temperature on solubility?

Ans: Effect of Temperature on solubility:

Temperature has major effect on the solubility of most of the substances. Generally it seems that solubility increases with the increase of temperature, but it is not always true. When a solution is formed by adding a salt in solvent there are three possibilities with reference to effect of temperature on solubility.

These possibilities are discussed here.

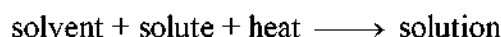
i. Heat is absorbed

ii. Heat is given out.

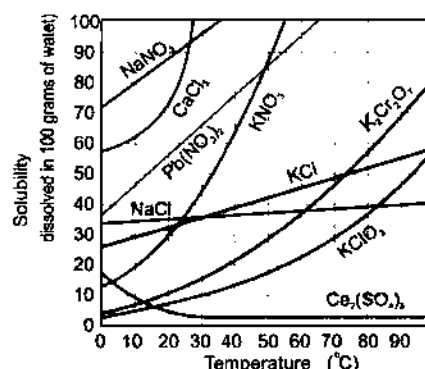
iii. No change in heat

Heat is absorbed (Endothermic Process)

Solubility usually increases with the increase in temperature for such solutes. When salts like KNO_3, $NaNO_3$ and KCl are added in water, the test tube becomes cold. It means during dissolution of these salts heat is absorbed. Such dissolving process is called "endothermic".

$$\text{solvent + solute + heat} \longrightarrow \text{solution}$$

Reason:

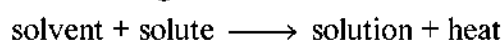
It means that heat is required to break the attractive forces between the ions of solute. This requirement is fulfilled by the surrounding molecules. As a result, their temperature falls down and test tube becomes cold.

Examples: KNO_3, $NaNO_3$, KCl, NH_4Cl, $CaCl_2$, $CuSO_4$ etc.

ii. Heat is given out (Exothermic Process).

Reason:

In such cases, the solubility of salt decreases with the increase of temperature

When salts like Li_2SO_4 and $Ce_2(SO_4)_3$ are dissolved in water, the test tube becomes warm. i.e. heat is released during this dissolution.

$$\text{solvent + solute} \longrightarrow \text{solution + heat}$$

In such cases attractive forces' among the solute particles are weaker and solute- solvent interactions are stronger. As a result, there is release of energy.

Examples: Li_2CO_3, $Ca(OH)_2$, Li_2SO_4,$CaCrO_4$ etc.

iii. No change in heat:

In some cases during a dissolution process neither the heat is absorbed nor released. When salt like NaCI is added in water, the solution temperature remains almost the same. In such case temperature has a minimum effect on solubility.

6.6 comparison of solution, suspension and colloid

Q:10 Explain the solution, suspension and colloid.

OR

a. Give the five characteristics of true solution.

b. (Ex.Q.8): Give the five characteristics of colloid.

c. (Ex.Q.9): Give at least five characteristics of suspension.

Ans: a. True Solution:

"A homogeneous mixture of two or more than two components is called true solution."

Examples: Solution of NaCl in water, drop of ink mixed in water (simplest example of true solution) and solution of sugar in water

Properties:

i. The particles exist in their simplest form i.e. as molecules or ions. Their diameter is 10^{-8}cm.

ii. Particles dissolve uniformly throughout and form a homogeneous mixture.

iii. Particles are so small that they can't be seen with naked eye.

iv. Solute particles can pass easily through a filter paper.

v. Particles are so small that they cannot scatter the rays of light, thus do not show

Tyndall effect.

b. Colloid / False solution / Colloidal Solution:

"These are solutions in which the solute particles are larger than those present in the true solutions but not large enough to be seen by naked eye."

Tyndall Effect:

"The particles of colloids are big enough to scatter the beam of light. It is called Tyndall effect."

Tyndall effect and distinction between colloid and solution:

We can see the path of scattered light beam inside the colloidal solution. Tyndall effect is the main characteristic which distinguishes colloids from solutions. Hence these solutions are called false solutions or colloidal solutions.

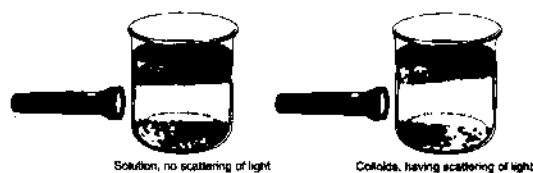



Fig. 6.4 Tyndall effect by colloids.

Examples: Starch, Albumin, Soap solutions, Blood, Milk, Ink, Jelly, Toothpaste etc

Properties:

i. The particles are large consisting of many atoms, ions or molecules.

ii. A colloid appears to be a homogeneous but actually it is a heterogeneous mixture. Hence, they are not true solution. Particles do not settle down for a long time, therefore, colloids are quite stable.

iii. Particles are large but can't be seen with naked eye.

iv. Although particles are big but they can pass through a filter paper.

v. Particles scatter the path of light rays thus emitting the beam of light i.e. exhibit the tyndall effect.

6.6.3 Suspension

"A heterogeneous mixture of undisclosed particles in a given medium that settles down after some time is called suspension."

Examples: Chalk in water (milky suspension), Paints and milk of magnesia (suspension of magnesium oxide in water)

Properties:

i. The particles are of largest size. They are larger than 10^{-5}cm in diameter.

ii. Particles remain un-dissolved and form a heterogeneous mixture. Particles settle down after sometime

iii. Particles are big enough to be seen with naked eye.

iv. Solute particles cannot pass through filter paper.

v. Particles are so big that light is blocked and difficult to pass.

Q:11 How you can compare solutions, colloid and suspension?

Ans: Comparison of the characteristics of solution, colloid and suspension

Solution	**Colloid**	**Suspension**
i. Size of particles: The particles exist in their simplest form i.e. as molecules or ions. Their diameter is 10^{-8} cm.	**i.** The particles are large consisting of many atoms, ions or molecules.	**i.** The particles are of largest size. They are larger than 10^{-5}cm in diameter.
ii. Solubility of particles: Particles dissolve uniformly throughout and form a homogeneous mixture.	**ii.** A colloid appears to be a homogeneous but actually it is a heterogeneous mixture. Hence, they are not true solution. Particles do not settle down for a long time, therefore, colloids are quite stable.	**ii.** Particles remain un-dissolved and Form a heterogeneous mixture. Particles settle down after sometime
iii. Observation with naked eye: Particles are so small that they can't be seen with naked eye.	**iii.** Particles are large but can't be seen with naked eye.	**iii.** Particles are big enough to be seen with naked eye.
iv. Passing through filter paper: Solute particles can pass easily through a filter paper.	**iv.** Although particles are big but they can pass through a filter paper.	**iv.** Solute particles cannot pass through filter paper.
v. Tyndall effect: Particles are so small that they cannot scatter the rays of light, thus do not show Tyndall effect.	**v.** Particles scatter the path of light rays thus emitting the beam of light i.e. exhibit the tyndall effect.	**v.** Particles are so big that light is blocked and difficult to pass.

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 6: Solutions
Extra MCQ's

1. **The emulsions are the type of solution:**
(a) Solid in solid (b) Liquid in solid (c) Gas in liquid (d) None of these
2. **In true solution, the particle are the type of solution:**
(a) 0.1mμ (b) 10.0 mμ (c) 0.5mμ (d) 1.0 mμ
3. **Which property is not for suspensions?**
(a) Homogeneous (b) Cannot be seen by naked eye
(c) Not a true solution (d) Solid in liquid
4. **The concentration of the solute in solution, when it is en equilibrium with solid substance, at a particular temperature is called:**
(a) Molarity (b) Dilution
(c) Colloidal solution (d) Supersaturated solution
5. **A solution containing relatively higher concentration of solute is called:**
(a) Dilute solution (b) Saturated solution (c) Concentrated solution (d) Suspension
6. **Fog is an example of solution:**
(a) Gas in liquid (b) Liquid in gas (c) Liquid in gas (d) Solid in liquid
7. **Emulsions are the colloidal dispersion of liquid in:**
(a) Solid (b) Gas (c) liquid (d) Water
8. **The suspension particles have appearance:**
(a) Uniform (b) Transparent (c) Opaque (d) Both 'A' and 'B'
9. **Starch, glue are the examples of:**
(a) Colloidal solution (b) Solution (c) Colloids (d) Crystalloids
10. **Solution which can dissolve further amount of a solute at particular temperature is called:**
(a) Saturated solution (b) Unsaturated solution
(c) Colloidal solution (d) Supersaturated solution
11. **The solution which can easily pass through parchment membrane is considered as:**
(a) Colloidal solution (b) True solution (c) Suspensions (d) Crystalloids
12. **Change of temperature can change the______ of a solute in a solvent.**
(a) Dilution (b) Solubility (c) Molarity (d) Dipole moment
13. **The ionic and polar compounds like NaCl and HCl are more soluble in water than non-polar covalent compounds like.**
(a) CCl_4 (b) Benzene (c) CS_2 (d) All of these
14. **Which one produced colloidal solution**
(a) Blood (b) Cooper Sulphate solution
(c) Silver nitrate solution (d) None of these
15. **Concentration is most often expressed as the ratio of the amount of _________to the amount of solution.**
(a) Solute (b) solvent (c) brine (d) salt
16. **Number of moles of solute per dm^3 of the solution is called:**
(a) Molarity (b) molality (c) normality (d) density
17. **Sodium amalgam which two metals as present**
(a) Na and Mg (b) Na and Mμ (c) Na and Hg (d) Zn and Co

18. The concentrated solution of NaCl is called:
(a) Fluid (b) Brass (c) Brine (d) Plasma

19. Brass is a familiar alloy, made of:
(a) zinc + carbon (b) Zinc + copper (c) copper + iron (d) zinc + lead

20. The homogeneous mixture of two or more compounds is called:
(a) solute (b) solvent (c) solution (d) fluid

21. Which of the following solutions contains least amount of water?
(a) 2M (b) 1M
(c) 0.5 M (d) 0.25M

22. Opal is an example of solution
(a) Liquid in solid (b) Solid in gas
(c) Solid in solid (d) Gas in liquid

23. 0.1M solution is diluted to ten time, its new morality is:
(a) 0.01M (b) 0.9M
(c) 0.2M (d) 0.1M

24. Which one of the following is true solution?
(a) Paints (b) Alcohol in water
(c) Milk (d) Milk of magnesia

25. Heat is absorbed on dissolving which one of the following salts:
(a) NaCl (b) $Ce_2(SO_4)_3$
(c) $NaNO_3$ (d) Li_2SO_4

26. Smoke is an example of solution
(a) Solid in gas (b) Gas in liquid
(c) Liquid in solid (d) Gas in solid

27. KCl is soluble in
(a) Water (b) Benzene
(c) Ether (d) Kerosene oil

28. Naphthalene is soluble in
(a) Water (b) Ether
(c) Carbon tetrachloride (d) Both (b) and (c)

29. Blood is an example of
(a) True solution (b) Suspension
(c) Colloid (d) None of all

30. 20 g of NaOH has been dissolved in $0.5dm^3$ of the solution its molarity is:
(a) 0.1M (b) 1.0M
(c) 0.5M (d) 1.5M

31. What mass of NaOH is required to prepare 1M $500cm^3$ of the solution:
(a) 10g (b) 20g
(c) 30g (d) 40g

32. What volume of 2M solution of H_2SO_4 is required to prepare 500 cm^3 0.1M solution?
(a) $10cm^3$ (b) $15cm^3$
(c) $20cm^3$ (d) $25cm^3$

33. 2g of NaCl has been dissolved in 20.0g of H_2O, its % m/m is
(a) 9.09% (b) 19.09%
(c) 1.9% (d) 10.0%

34. Identify among the following which one is universal solvent?
(a) CH_3COOH (b) Liquid ammonia
(c) Water (d) Ethanol

35. Which one of the following salts gives out heat on dissolving in water?
(a) NaCl (b) $Ce_2(SO_4)_3$

(c) KNO_3 (d) KCl

36. Mist is an example of solution

(a) Liquid in gas (b) Gas in liquid
(c) Solid in gas (d) Gas in solid

37. Which one of the following is a liquid in solid solution?

(a) Sugar in water (b) Butter
(c) Solvent to solution (d) Fog

38. Which one of the following solutions contains more water?

(a) 2 M (b) 1 M
(c) 0.5 M (d) 0.25 M

39. Tynadall effect is shown by

(a) Sugar solution (b) Jelly
(c) Paints (d) Chalk solution

40. Tyndall effect is due to

(a) Blockage of beam of light (b) Non-scattering of beam of light
(c) Scattering of beam of light (d) Passing through beam of light

41. If $10cm^3$ of alcohol is dissolved in 100g of water, it is called

(a) % m/m (b) % m/v
(c) % v/m (d) % v/v

42. Sea water is a source of

(a) Fog (b) Opal
(c) Soft drink (d) Cheese

43. Which one of the following solutions contains maximum amount of water?

(a) 2M (b) 1M
(c) 0.5 M (d) 0.25 M

44. Brass is a solid solution of

(a) Zn and Cd (b) Zn and Cu
(c) Zn and Hg (d) Zn and Mg

45. $10cm^3$ of 0.01 molar $KMnO_4$ solution has been diluted to 100 cm^3 its molarity will become

(a) 0.001 M (b) 0.1 M
(c) 0.01M (d) 1 M

ANSWER KEY

1	b	11	b	21	a	31	b	41	c
2	a	12	b	22	c	32	d	42	b
3	b	13	d	23	a	33	a	43	d
4	d	14	a	24	b	34	c	44	b
5	c	15	a	25	c	35	b	45	a
6	b	16	b	26	a	36	a		
7	a	17	c	27	a	37	d		
8	c	18	c	28	d	38	d		
9	c	19	b	29	c	39	b		
10	b	20	c	30	b	40	c		

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 7: Electrochemistry
Exercise Questions

Exercise Multiple Choice Question Answers:

1. **Spontaneous chemical reactions take place in:**
   (a) Electrolytic cell (b) Galvanic cell (c) Nelson's cell, (d) Down's cell
2. **Formation of water from hydrogen and oxygen is:**
   (a) Redox reaction (b) Acid-base reaction
   (c) Neutralization (d) Decomposition
3. **Which one of the following is not an electrolytic cell?**
   (a) Downs cell (b) Galvanic cell (c) Nelson's cell (d) Both a and c
4. **The oxidation number of chromium in $K_2Cr_2O_7$ is:**
   (a) +2 (b) +6 (c) +7 (d) +14
5. **Which one of the following is not an electrolyte?**
   (a) Sugar solution (b) Sulphuric acid solution
   (c) Lime solution (d) Sodium chloride solution
6. **The most common example of corrosion is:**
   (a) Chemical decay (b) Rusting of iron
   (c) Rusting of aluminium (d) Rusting of tin
7. **Nelson's cell is used to prepare caustic soda along with gases. Which of the following gas is produced at cathode:**
   (a) Cl_2 (b) H_2 (c) O_3 (d) O_2
8. **During the formation of water from hydrogen and oxygen, which of the following does not occur:**
   (a) Hydrogen has oxidized (b) Oxygen has reduced
   (c) Oxygen gains electrons (d) Hydrogen behaves as oxidizing agent
9. **The formula of rust is:**
   (a) $Fe_2O_3.nH_2O$ (b) Fe_2O_3 (c) $Fe(OH)_3.nH_2O$ (d) $Fe(OH)_3$
10. **In the redox reaction between Zn and HCl, the oxidizing agent is:**
    (a) Zn (b) H^+ (c) Cl^- (d) H

ANSWR KEY

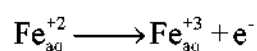
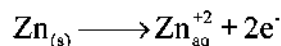
1	b	2	a	3	d	4	b	5	a	6	b	7	b
8	d	9	a	10	b								

Exercise Short Question Answers

Q.1 Define oxidation in terms of electrons. Give an example.

Ans: Oxidation is the loss of electron by an atom or an ion e.g,

$$Zn_{(s)} \longrightarrow Zn^{+2}_{aq} + 2e^-$$

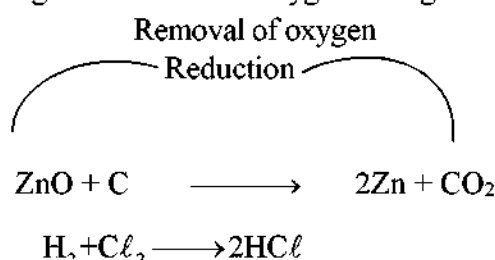
$$Fe^{+2}_{aq} \longrightarrow Fe^{+3}_{aq} + e^-$$

Q.2 Define reduction in terms of loss or gain of oxygen or hydrogen. Give an example.

Ans: Reduction:

"The addition of hydrogen or removal of oxygen during a chemical reaction."

Examples:

i. $ZnO + C \longrightarrow 2Zn + CO_2$

ii. $H_2 + C\ell_2 \longrightarrow 2HC\ell$

Q.3 What is difference between valency and oxidation state?

Ans:

Valency	Oxidation Number or state
• The combining capacity of an element with other element is called value of	• The apparent charge assigned to an atom of an element in a molecule or ion is called oxidation state.
• While assigning valency the sign is followed by the number i,e 2+	• No sign
• For example valency of sodium is 1+	• For example oxidation number of sodium is +1

Q.4 Differentiate between oxidizing and reducing agents

Ans:

Oxidizing agent	Reducing Agent
i. A species that oxidizes a substance by taking electrons from it, is called an oxidizing agent.	**i.** A species that reduces a substance by donating electrons to it is called reducing agent.
ii. Non metals are good oxidizing agents.	**ii.** Metals are good reducing agents.
iii. They are more electronegative in nature.	**iii.** They are more electropositive.
iv. Its oxidation number decreases.	**iv.** Its oxidation number decreases.
v. Examples	**vii.** Examples
vi. $S + O_2 \longrightarrow SO_2$	**viii.**

Q.5 Differentiate between strong and weak electrolytes.

Ans:

Strong electrolyte	Weak electrolyte
The electrolyte which ionize completely in solution and produce more ions, are called strong electrolyte.	*The electrolytes which ionize to a small extent when dissolve in water and could not produce more ions are called weak electrolytes.*
Examples: NaCl, NaOH, H_2SO_4 etc.	Examples: $Ca(OH)_2$, CH_3COOH $CH_3COOH_{\ell} + H_2O_{\ell} \longrightarrow CH_3COO^-_{aq} + H_3O^+$

Q.6 How electroplating of tin on steel is carried out?

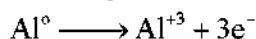
Ans: In electroplating of silver, when current is passed through the cell. A.g. ions present in the electrolyte solution migrate towards the cathode and deposit after picking up electrons. The anode consists of silver bar or sheet. Which is oxidized to Ag ions which dissolve in solution and migrate towards the cathode where they are discharged and deposited on the object

At anode: $Ag_{(s)} \longrightarrow Ag^+_{(aq)} + \bar{e}$

At cathode: $Ag^+_{(aq)} \longrightarrow Ag_{(s)}$

Q.7 Why steel is plated with nickel before the electroplating of chromium.

Ans: The steel is usually plated first with nickel or copper then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is stripped off.

Q.8 How can you explain, that following reaction is oxidation in terms of increase of oxidation number Al" $Al^o \longrightarrow Al^{+3} + 3e^-$

Ans: Increase in oxidation number is called oxidation oxidation number of Al in creases from zero to + 3 as given below $A\ell \longrightarrow A\ell^{+3} + 3e^-$

Q.9 How can you prove so it is an oxidation reaction with an example that conversion of an ion to an atom is an oxidation process?

Ans: Conversion of anion into an atom is an oxidation process.

Example:

When anions (negatively charged ions) lose electron, they are converted into atoms and oxidized.

$$Cl^- \xrightarrow{Oxidation} Cl + le^-$$

Q.10 Why does the anode carries negative charge in galvanic cell but positive charge in electrolytic cell? Justify with comments.

Ans: In Gavanic cell, electrons are lost by the atoms at anode plate which makes it electron efficient therefore it carries negative charge. In electrolytic cell, electrons are gained by cations from anode which makes it electron deficient therefore it carries positive charge.

Q.11 Where do the electrons flow from Zn electrode in Daniel's cell?

Ans: In Daniel cell, the electrons takes flow from Zn electrode (anode) towards the cathode made up of copper through the external circuit.

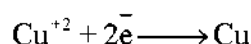
Q.12 Why do electrodes get their names 'anode' and cathode in galvanic cell?

Ans: In galvanic cell anode and cathode get their names depending upon the process taking place on them.

Anode: is an electrode where oxidation takes place

e.g. $$Zn \longrightarrow Zn^{+2} + 2e^-$$

Cathode: is an electrode where reduction takes place

$$Cu^{+2} + 2\bar{e} \longrightarrow Cu$$

In galvanic cell, oxidation takes place at anode while reduction takes place at cathode. And oxidation always takes place at anode while reduction always takes place at cathode.

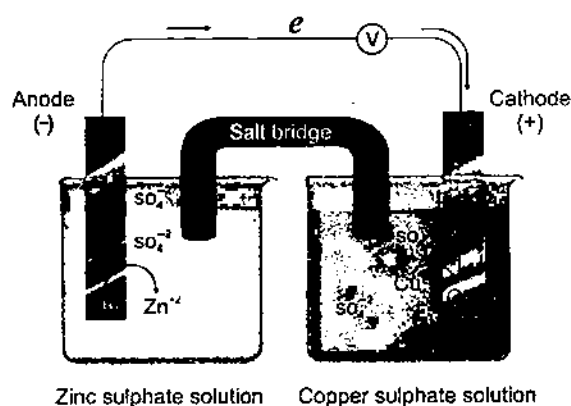
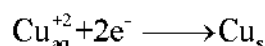



Fig. 7.3 A Daniel Cell

Q.13 What happens at the cathode in a galvanic cell?

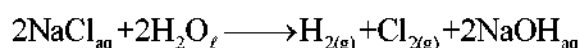
Ans: In galvanic cell, reduction takes place at the cathode as:

$$Cu^{+2}_{aq} + 2e^- \longrightarrow Cu_s$$

Q.14 Which solution is used as an electrolyte in Nelson's cell?

Ans: An (aqueous solution of NaCl called brine? is used as electrolyte in Nelson's cell.

Q.15 Name the by-products produced in Nelson's cell?

Ans: Hydrogen gas (H_2) and chlorine gas (Cl_2) are the by-product of Nelson's cell as

$$2NaCl_{aq} + 2H_2O_{\ell} \longrightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{aq}$$

Q.16 Why galvanizing is done?

Ans: The process of coating a thin layer of zinc on iron is called galvanizing. Galvanizing is done to protect the iron against corrosion even after the required coating surface is broken.

Q.17 Why an iron grill is painted frequently?

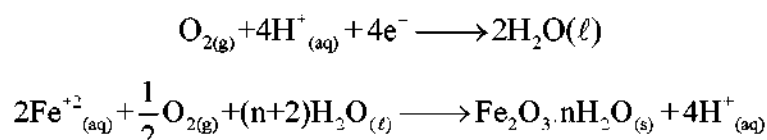
Ans: Iron grill is painted frequently to protect it from rusting. Paint layer proctect iron from attack of moisture and oxygen.

Q.18 Why O_2 is necessary for rusting?

Ans: O_2 is necessing for rusing because it acts as oxidizing agent. It accepts electrons from Fe which is covered to Fe+2 and then to Fe+3. Oxygen combines with Fe+3 to form rust (Fe_2O_3 H_2O)

The overall cell nraction for corrosion of ions is

$$O_{2(g)} + 4H^{+}_{(aq)} + 4e^{-} \longrightarrow 2H_2O(\ell)$$

$$2Fe^{+2}_{(aq)} + \frac{1}{2}O_{2(g)} + (n+2)H_2O_{(\ell)} \longrightarrow Fe_2O_3.nH_2O_{(s)} + 4H^{+}_{(aq)}$$

Q.19 In electroplating of chromium, which salt is used as an electrolyte?

Ans: Chromium sulphate with few drops of H_2SO_4 acts as electrolyte.

Q.20 Write the redox reaction taking place during the electroplating of chromium?

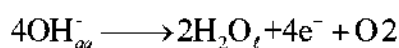
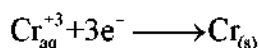
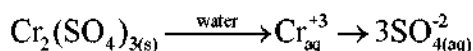
Ans: At anode:

$$4OH^{-}_{aq} \longrightarrow 2H_2O_{\ell} + 4e^{-} + O2$$

At cathode:

$$Cr^{+3}_{aq} + 3e^{-} \longrightarrow Cr_{(s)}$$

Overall reaction:

$$Cr_2(SO_4)_{3(s)} \xrightarrow{water} Cr^{+3}_{aq} \rightarrow 3SO^{-2}_{4(aq)}$$

Q.21 In electroplating of silver, from where Ag + come and where they deposit?

Ans: In electroplating of silver Ag^{+} ion come form anode while they deposit at cathode.

Q.22 What is the nature of electrode used in electroplating of chromium?

Ans: In electroplating of chromium, anode is made of antimonial lead while the object to be electroplated acts as cathode.

Exercise Long Question Answers

Q.1 Describe the rules for assigning the oxidation state

Ans: See Q. No. **4** (Subjective Part, Long Questions Answers)

Q.2 Find out the oxidation numbers of the underlined elements in the following compounds.

(a) $Na_2S_2O_4$

(b) $K_2Cr_2O_7$

(c) $AgNO_3$

(d) HNO_3

(e) $KMnO_4$

Ans:

Q.3 How can a non-spontaneous reaction be carried out in an electrolytic cell. Discuss in detail.

Ans: See Q. No. **7** (Subjective Part, Long Questions Answers)

Q.4 Discuss the electrolysis of water.

Ans: See Q. No. **8** (Subjective Part, Long Questions Answers)

Q.5 Discuss the construction and working of a cell in which electricity is produced.

Ans: See Q. No. **9** (Subjective Part, Long Questions Answers)

Q.6 How we can prepare NaOH on commercial scale. Discuss its chemistry along with the diagram.

Ans: See Q. No. **12** (Subjective Part, Long Questions Answers)

Q.7 Discuss the redox reaction taking place in the rusting of iron in detail.

Ans: See Q. No. **13** (Subjective Part, Long Questions Answers)

Q.8 Discuss, why galvanizing is considered better than that of tin plating.

Ans: See Q. No. **14** (Subjective Part, Long Questions Answers)

Q.9 What is electroplating? Write down procedure of electroplating.

Ans: See Q. No. **15** (Subjective Part, Long Questions Answers)

Q.10 What is the principle of electroplating? How electroplating of chromium is carried out?

Ans: See Q. No. **16** (Subjective Part, Long Questions Answers)

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 7: Electrochemistry
Long Questions

Q.1 a. Define electrochemistry

b. Differentiate between spontaneous and non-spontaneous reactions

Ans: a .Electrochemistry

The branch of Chemistry that deals with the relationship between electricity and chemical reactions is called electrochemistry.

OR

The branch of chemistry that deals with the study of conversion of chemical energy into electrical energy and electrical energy into chemical energy is called electrochemistry.

Electrochemistry involves oxidation and reduction reactions, which are also known as redox reactions. There are two types of redox reactions.

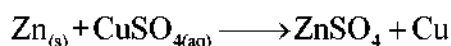
i. Spontaneous reaction **ii.** Non-spontaneous reactions

b. Difference between spontaneous and non-spontaneous reactions:

i. Spontaneous reactions

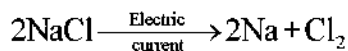
Spontaneous reactions are those which take place on their own without an external agent.

For example, reactions taking place in voltaic or galvanic cells and corrosion process

$$Zn_{(s)} + CuSO_{4(aq)} \longrightarrow ZnSO_4 + Cu$$

ii. Non-spontaneous reactions

Non-spontaneous reactions are those which take place in the presence of an external agent.

For example, reactions taking place in electrolytic cells like electrolysis of fused NaCl to produce sodium.

$$2NaCl \xrightarrow[\text{current}]{\text{Electric}} 2Na + Cl_2$$

Q.2 Explain Oxidation and Reduction reactions on the basis of addition and removed of hydrogen and oxygen.

Ans: Oxidation and reduction reactions in term of loss or gain of hydrogen/oxygen:

Oxidation:

Oxidation is defined as addition of oxygen or removal of hydrogen during a chemical reaction.

Reduction:

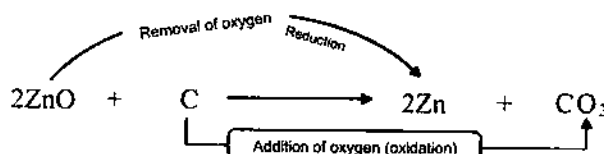
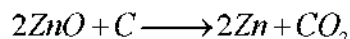
Reduction is defined as addition of hydrogen or removal of oxygen during a chemical reaction.

Both of these processes take place simultaneously in a redox reaction.

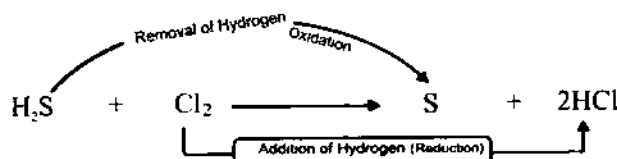
Examples:

i. Addition or removal of oxygen:

A reaction between zinc oxide and carbon takes place by the removal of oxygen (reduction) from zinc oxide and addition of oxygen (oxidation) to carbon.

$$2ZnO + C \longrightarrow 2Zn + CO_2$$

ii. Addition or removal of hydrogen:

A reaction between hydrogen sulphide and chlorine takes place by oxidation of hydrogen sulphide and reduction of chlorine. Hydrogen is being removed from H_2S and added to chlorine.

$$H_2S + Cl_2 \longrightarrow HCl + S$$

Q.3 Explain oxidation and reduction in terms of loss or gain of electrons.

Ans: Oxidation and reduction in terms of loss or gain of electrons.

There are many chemical reactions which do not involve oxygen or hydrogen but they are considered redox reactions. To deal with these reactions, a new concept "loss or gain of electrons" is used called electronic concept.

According to this concept:

Oxidation:

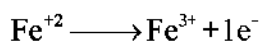
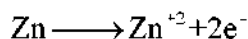
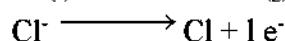
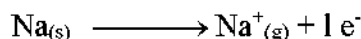
Oxidation is loss of electrons by an atom or an ion.

e.g.

$$Na_{(s)} \longrightarrow Na^{+}_{(g)} + 1e^-$$

$$Cl^- \longrightarrow Cl + 1e^-$$

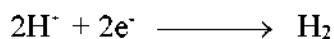
$$Zn \longrightarrow Zn^{+2} + 2e^-$$

$$Fe^{+2} \longrightarrow Fe^{3+} + 1e^-$$

Reduction:

Reduction is gain of electrons by an atom or ion.

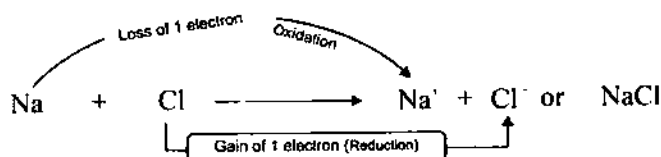
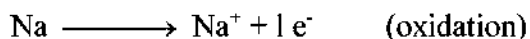
e.g.

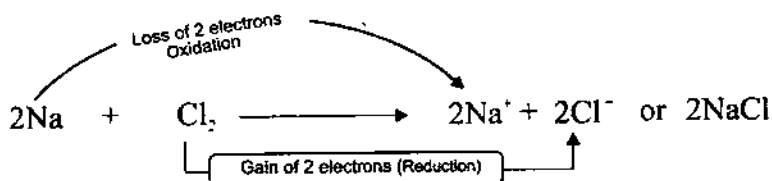
$$2H^+ + 2e^- \longrightarrow H_2$$

$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$

Explanation:

A reaction between sodium metal and chlorine takes place in three steps.

i. First sodium atom losses an electron, to form sodium ion.

$$Na \longrightarrow Na^+ + 1\,e^- \quad \text{(oxidation)}$$

ii. Simultaneously, this electron is accepted by chlorine atom (reduction process), as chlorine atom needs one electron to complete its octet. As a result chlorine atom changes to chloride ion.

$$Cl\,(g) + 1e^- \longrightarrow Cl^- \text{ (reduction)}$$

iii. Ultimately, both these ions attract each other to form sodium chloride. Complete redox reaction is sum of the oxidation and reduction reactions between sodium and chlorine atoms and it is represented as;

$$Na^+ + Cl^- \longrightarrow NaCl$$

Summary of concept of oxidation and reduction

Oxidation	Reduction
Addition of oxygen	*Removal of oxygen*
Removal of hydrogen	*Addition of hydrogen*
Loss of electrons	*Gain of electrons*

Q.4 Describe the rules for assigning oxidation states. (Ex.Q.1)

Oxidation states:

Oxidation state or oxidation number (O.N.) is the apparent charge assigned to an atom of an element in a molecule or in an ion.

It may be positive or negative or whole number or in fraction or zero.

Examples:

In HCl, oxidation number of H is +1 and that of Cl is –1.

Rules for assigning oxidation numbers (O.N):

i. O.N of elements in free state:

The oxidation number of all elements in the free state is zero.

ii. O. N. of simple ions:

The oxidation number of an ion consisting of a single element is the same as the charge on the ion.

iii. O. N of elements in complex ions:

The oxidation number of different elements in Group-I is +1, in Group-2 is +2 and in Group-3 is +3.

iv. **O.N of H:**
The oxidation number of hydrogen in all its compounds is + 1 but in metal hydrides it is -1.

v. **O. N. of oxygen:**
The oxidation number of oxygen in all its compounds is -2 but it is -1 in peroxides and +2 in OF_2.

vi. **Negative O.N:**
In any substance the more electronegative atom has the negative oxidation number.

vii. **O.N. of neutral molecules:**
In neutral molecules, the algebraic sum of the oxidation numbers of all the elements is zero.

viii. **O.N. in ions:**
In ions, the algebraic sum of oxidation number equals the charge on the ion.

Q.5 Explain oxidizing and reducing agents with the help of suitable examples.

Ans: Oxidizing agents:

An oxidizing agent is the species that oxidizes a substance by taking electrons from it.

OR

The substance (atom or ion) which is reduced itself by gaining electrons from other substance is also called oxidizing agent.

Examples: Non-metals $(O_2, C\ell_2, F_2)$ are oxidizing agents because they accept electrons being more electronegative elements. Strong acids like HNO_3, H_2SO_4 etc are also metals.

Reducing agents:

Reducing agent is the species that reduces a substance by donating electron to it.

OR

The substance (atom or ion) which is oxidized by losing electrons to other substance is also called reducing agent.

Examples: Metals like zinc, iron, aluminum etc. are good reducing agent.

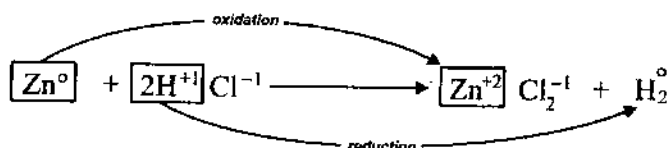
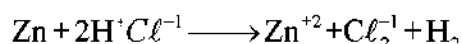
Explanation:

i. Reaction between Zn and $HC\ell$**:**

Let us discuss a reaction of Zn metal with hydrochloric acid.

$$Zn_{(s)} + HC\ell_{(aq)} \longrightarrow ZnC\ell_2 + H_{2(g)}.$$

The oxidation states or oxidation numbers of all the atoms or ions in this reaction are indicated below

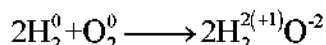
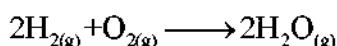
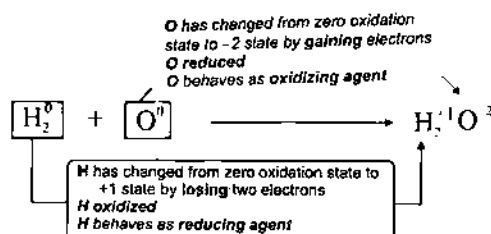
$$Zn + 2H^{+}C\ell^{-1} \longrightarrow Zn^{+2} + C\ell_2^{-1} + H_2$$

Let us find the atoms, which are oxidized or reduced or whether there is a change in their oxidation state. It is indicated as follows.

In this reaction, zinc is oxidized and acts as a reducing agent. Hydrogen is reduced and acts as an oxidizing agent.

ii. Reaction between H & O:

In case of formation of water from hydrogen and oxygen gases, redox reaction takes place as follows.

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$$

$$2H_2^0 + O_2^0 \longrightarrow 2H_2^{2(+1)}O^{-2}$$

In this reaction, H is oxidized and acts as are reducing agent. O is reduced and acts as an oxidizing agent.

Q.6 What are electrochemical cells? Explain the concept of electrolytes.

Ans: Electrochemical cell:

Electrochemical cell is a system in which two electrodes are dipped in the solution of an electrolyte or molten mass.

Electrochemical cell is an energy storage device in which either a chemical reaction takes place by using electric current (electrolysis) or chemical reaction produces electric current.

Types of electrochemical cells:

i. Electrolytic cells ii. Galvanic cells or Voltaic cells

i. Electrolytic cells:

The type of electrochemical reaction in which a non-spontaneous chemical reaction takes place when electric current is passed through an electrolyte is called an electrolytic cell.

Examples: Down's cell, Nelson cell etc.

ii. Galvanic cells or Voltaic cells:

The type of electrochemical cell in which a spontaneous chemical reaction takes place and generates electric current is called Galvanic or Voltaic cell.

Example: Daniel cell, Dry cell.

Concept of Electrolytes:

a. Electrolytes

The substances, which can conduct electricity in their solutions or molten states, are called electrolytes.

Solutions of salts, acids or bases are good electrolytes. The electricity cannot pass through solid NaCl but its aqueous solution or molten NaCl are good electrolytes.

Classification of electrolytes:

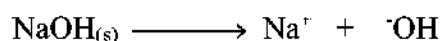
Electrolytes are classified into two groups depending upon their extent of ionization in solution.

i. Strong Electrolytes

The electrolytes which ionize completely in solution and produce more ions, are called

strong electrolytes.

Examples: Strong electrolytes are aqueous solutions of NaCl, NaOH and H_2SO_4 etc.

$$NaOH_{(s)} \longrightarrow Na^{+} + {}^{-}OH$$

ii. Weak electrolytes:

The substances which ionize to a small extent when dissolved in water and produce less ions are called weak electrolytes.

Examples: Acetic acid (CH_3COOH), and $Ca(OH)_2$ when dissolved in water, ionizes to a small extent. These are good examples of weak electrolytes. Weak electrolytes do not ionize completely. For example, ionization of acetic acid in water produces less ions:

$$CH_3COOH_{(l)} + H_2O_{(l)} \longrightarrow CH_3COO^{-}_{(aq)} + H_3O^{+}_{(aq)}$$

As a result the weak electrolyte is a poor conductor of electricity.

b. Non-Electrolytes

The substances, which do not ionize in solution and do not allow the current to pass through their solutions, are called non-electrolytes.

Example: Sugar solution and benzene are non-electrolytes.

Q.7 (Ex. Q.s)What is electrolysis? Writes a note on electrolytic cells.

OR

How can a non- spontaneous reaction be carried out in on electrolytic cell? Discuss in detail.

Ans: Electrolysis:

The chemical decomposition of a compound into its components by passing current through the solution of the compound or in the molten state of the compound is called electrolysis.

Example: Electrolyses of NaCl into Na and Cl.

Electrolytic Cell:

The type of electrochemical cell in which a non-spontaneous chemical reaction takes place when electric current is passed through the electrolyte, is called an electrolytic cell.

Example: Down's cell, Nelson's cell, Hoffmann's Voltameter.

Construction of an Electrolytic Cell:

An electrolytic cell consists of a solution of an electrolyte, two electrodes (anode and cathode) that are dipped in the electrolytic solution and connected to the battery.

Anode:

The electrode connected to positive terminal is called anode

Cathode:

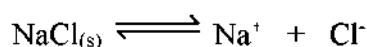
Electrode connected to the negative terminal is called cathode.

Working of an Electrolytic Cell:

When electric current is applied from battery, the ions in the electrolyte migrate to their respective electrodes. The anions, which are negatively charged, move towards the anode and discharge there by losing their electrons. Thus oxidation takes place at anode. While cations, which are positively charged ions, move towards cathode. Cations gain electrons from the cathode and as a result reduction takes place at cathode.

Example (Electrolysis of fused sodium chloride):

When fused salt of sodium chloride is electrolyzed the following reactions take place during this process.

Ionization reaction:

$$NaCl_{(s)} \rightleftharpoons Na^+ + Cl^-$$

Oxidation reaction at anode:

$$2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^-$$

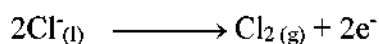
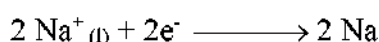
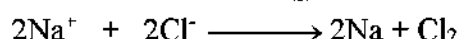
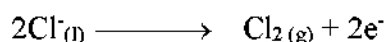
Reduction reaction at cathode:

$$2\,Na^+{}_{(l)} + 2e^- \longrightarrow 2\,Na$$

Overall reaction:

$$2Na^+ + 2e^- \longrightarrow 2Na$$

$$2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^-$$

$$2Na^+ + 2Cl^- \longrightarrow 2Na + Cl_2$$

Q.8 Discuss the electrolysis of water.

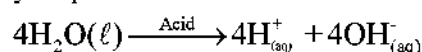
Ans: Electrolysis:

The chemical decomposition of a compound into its components by passing current through the solution of the compound or in the molten state of the compound is called electrolysis.

Electrolysis of water:

Extent of ionization of water:

Pure water is a very weak electrolyte. It ionizes to a very small extent. The concentrations of hydrogen ions (H^+) and hydroxide ions (^-OH) are both at 10^{-7} mol dm^{-3} respectively. Increase in conductivity of water. When a few drops of an acid are added in water, its conductivity improves.

$$4H_2O(\ell) \xrightarrow{Acid} 4H^+_{(aq)} + 4OH^-_{(aq)}$$

Working:

When an electric current is passed through this acidified water, OH^- **(anions)** move towards positive electrode (anode) and H^+ **(cations)** move towards negative electrode (cathode) and discharge takes place at these electrodes. They produce oxygen and hydrogen gases respectively at anode and cathode.

Oxidation reaction at anode:

$$4OH^-_{(aq)} \longrightarrow 2H_2O_{(\ell)} + O_{2(g)} + 4e^-$$

Reduction reaction at cathode:

$$4H^+_{(aq)} + 4e^- \longrightarrow 4H_{(g)}$$

$$4H_{(g)} \longrightarrow 2H_{2(g)}$$

Overall reaction:

$$2H_2O_{(\ell)} \longrightarrow O_{2(g)} + 2H_{2(g)}$$

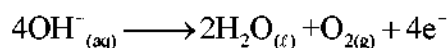
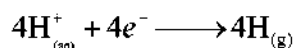
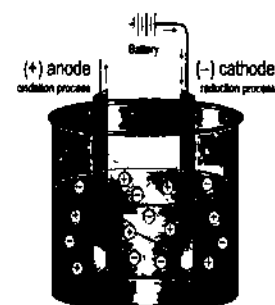
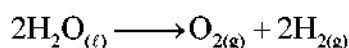
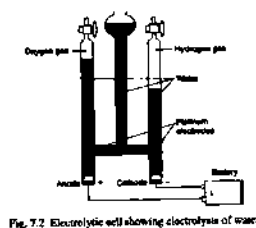



Fig. 7.2 Electrolytic cell showing electrolysis of water

Q.9 Discuss the construction and working of a cell in which electricity is produced.

OR

What is meant by Galvanic cell? Write construction and working of Daniel's cell.

Ans: Galvanic cell:

The electrochemical cell in which a spontaneous chemical reaction takes place and generates electric current is called Galvanic or Voltaic cell.

Examples: Daniel cell, Dry cell etc.

Introduction:

A. Volta (1745-1827) was an Italian physicist known especially for the development of the first electric cell in 1800.

Construction and working of Daniel cell:

a. Construction: galvanic cell consists of two cells, each called as half cell, connected electrically by a salt-bridge. In each of the half-cell, an electrode is dipped in 1M solution of its own salt and connected through a wire to an external circuit.

i. Left half cell (Oxidation half-cell):

The left half cell consists of an electrode of zinc metal dipped in 1 M solution of zinc sulphate.

ii. Right half Cell (Reduction half cell):

The right half cell is a copper electrode dipped in 1M solution of copper sulphate.

iii. Salt bridge:

Salt bridge is a U shaped glass tube. It consists of saturated solution of strong electrolyte supported in a jelly type material. The ends of the U tube are sealed with a porous material like glass wool.

Function of the salt bridge:

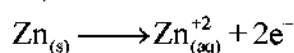
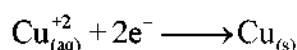
To keep the solutions of two half cells neutral by providing a pathway for migration of Ions.

b. Working of the Cell: The Zn metal has tendency to lose electrons more readily than copper. As a result oxidation takes place at Zn-electrode. The electrons flow from Zn-electrode through the external wire in a circuit to copper electrode. These electrons are gained by the copper ions of the solution and copper atoms deposit at the electrode. The respective oxidation and reduction processes going on at two electrodes are as follows:

Reaction at anode (oxidation):

$$Zn_{(s)} \longrightarrow Zn^{+2}_{(aq)} + 2e^-$$

Reaction at cathode (reduction)

$$Cu^{+2}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$$

Overall galvanic reaction is the sum of these two half cell reactions:-

$$Zn_{(s)} + CU^{+2}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Cu_{(s)}$$

As a result of redox reaction electric current is produced.

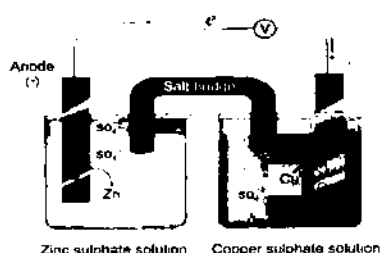



Fig. 7.3 A Daniel Cell

Uses:

The batteries which are used for starting automobiles, running calculators and toys etc. work on the principle of Galvanic cell.

Q.10 Differentiate between electrolytic and galvanic cell

Ans: The Comparison of Electrolytic and Galvanic Cells

Electrolytic cell	Galvanic cell
i. It consists of one complete cell, connected to a battery.	**i.** It consist of two half cells connected, through a salt bridge.
ii. Anode has positive charge while cathode has negative charge.	**ii.** Anode has negative charge while cathode has positive charge.
iii. Electrical energy is converted into chemical energy.	**iii.** Chemical energy converted into electrical energy.
iv. Current is use for a non-spontaneous chemical reaction to take a place.	**iv.** Redox reaction takes place spontaneously and produce electric current.
Examples: Down's cell, Nelson cell	**Examples:** Daniel cell, dry Cell etc.

Q.11 How sodium metal is manufactured from fused NaCl?

Ans: Manufacture of Sodium Metal from Fused NaCl:

Principle:

On the industrial scale molten sodium metal is obtained by the electrolysis of fused NaCl in the Downs cell. Construction of Down's cell.

Construction and working of Down's cell:

a. Construction

This electrolytic cell is a circular furnace. In the center there is a large block of graphite, which acts as an anode while cathode around it is made of iron.

b. Working:

i. Ionization:

The fused NaCl produces Na^+ and Cl^- ions, which migrate to their respective electrodes on passage electric current.

ii. Separation of electrodes:

The electrodes are separated by steel gauze to prevent the contact between the products.

iii. Reactions:

The Cl^- ions are oxidized to give Cl_2 gas at the anode. It is collected over the anode within an inverted cone-shaped structure while Na^+ are reduced at cathode and molten Na metal floats on the denser molten salt mixture from where it is collected in a side tube. Following reactions take place during the electrolysis of the molten sodium chloride:

Downs Cell for production of Sodium Metal

Ionization of molten NaCl

$$2NaCl \longrightarrow 2Na^+ + 2Cl^-$$

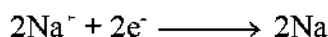
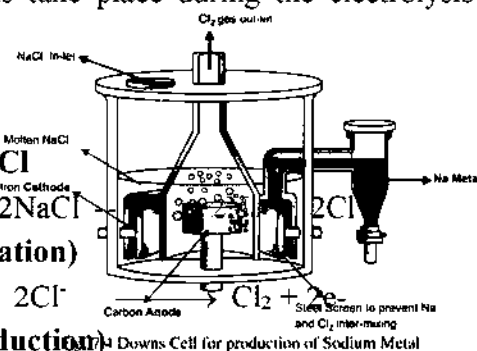
Reaction at anode (oxidation)

$$2Cl^- \longrightarrow Cl_2 + 2e^-$$

Reaction at cathode (reduction)

$$2Na^+ + 2e^- \longrightarrow 2Na$$

Overall reaction

$$2NaCl \longrightarrow 2Na + Cl_2$$

Q.12 How can we prepare NaOH on commercial scale? Discuss its chemistry along with diagram?

Ans: Manufacture of NaOH (Caustic soda):

Principle:

On industrial scale caustic soda, sodium hydroxide NaOH, is produced in Nelson's cell by the electrolysis of aqueous solution of NaCl, called brine.

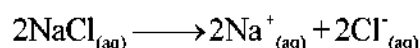
Construction and working of Nelson's cell:

a. Construction: It consists of a steel tank in which graphite anode is suspended in the center of a U shaped perforated iron cathode. This iron cathode is internally lined with asbestos diaphragm. Electrolyte (brine) is present inside the iron cathode.

b. Working:

i. Ions Present:

Aqueous solution of sodium chloride consists of Na^+, Cl^-, H^+ and ^-OH ions. These ions move towards their respective electrodes and redox reactions take place at these electrodes. When electrolysis takes place Cl ions are discharged at anode and Cl_2 gas rises into the dome at the top of the cell. The H^+ ions are discharged at cathode and H_2 gas escapes through a pipe. The sodium hydroxide solution slowly percolates into a catch basin.

Ionization of Brine:

$$2NaCl_{(aq)} \longrightarrow 2Na^+_{(aq)} + 2Cl^-_{(aq)}$$

Reaction at anode (oxidation):

$$2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e^-$$

Reaction at cathode (reduction):

$$2H_2O + 2e^- \longrightarrow H_{2(g)} + 2OH^-_{(aq)}$$

$$2Na^+_{(aq)} + 2OH^-_{(aq)} \longrightarrow 2NaOH_{(aq)}$$

Overall cell reaction:

$$2NaCl_{(aq)} + 2H_2O_{(\ell)} \longrightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{(aq)}$$

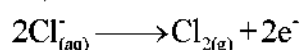
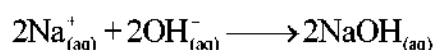
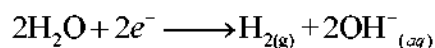
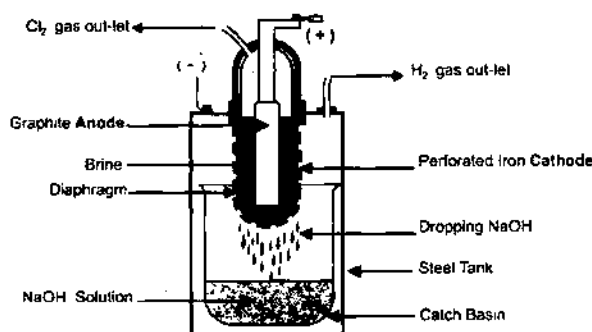
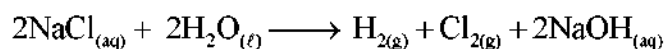



Fig. 7.5 Nelson's Cell for production of NaOH

Q.13 What is Corrosion? How iron gets rusted?

Ans: Corrosion:

Corrosion is slow and continuous eating away of a metal by the surrounding medium. Corrosion is a redox chemical reaction that takes place by the action of air and moisture with the metals. The **most common example** of corrosion is **rusting of iron.**

Rusting of Iron:

Corrosion is a general term but corrosion of iron is called rusting. Formation of hydrated iron oxide $(Fe_2O_3.nH_2O)$ *at surface of iron is called rusting.*

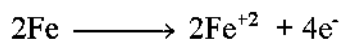
Conditions for rusting:

The important condition for rusting is moist air (air having water vapours in it). There will be no rusting in water vapours free of air or air free of water.

Process of rusting:

i. Anodic region:

Stains and dents on the surface of the iron provide the sites for this process to occur. This region is called anodic region and following oxidation reaction takes place here:

$$2Fe \longrightarrow 2Fe^{+2} + 4e^-$$

This loss of electrons damages the object.

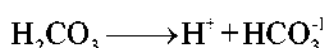
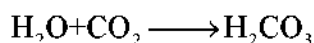
ii. Cathodic region

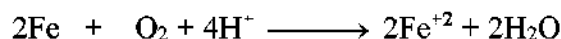
The free electrons move through iron sheet ,until they reach to a region of relatively high O_2 concentration near the surface surrounded by water layer. This region acts as cathode and electrons reduce the oxygen molecule in the presence of H+ ions:

Provision of H^+ ions:

The H^+ ions are provided by the carbonic acid, which are formed because of presence of CO_2 in water. That's why acidic medium accelerates the process of rusting

$$H_2O + CO_2 \longrightarrow H_2CO_3$$

$$H_2CO_3 \longrightarrow H^+ + HCO_3^{-1}$$

The overall redox process is completed without the formation of rust

$$2Fe + O_2 + 4H^+ \longrightarrow 2Fe^{+2} + 2H_2O$$

Formation of rest:

The Fe^{+2} formed spreads through out the surrounding water and react with O_2 to form the salt $Fe_2O_3 .nH_2O$ which is called **rust.** It is also a redox reaction.

The rust layer of iron is porous and does not prevent further corrosion. Thus rusting continues until all the piece of iron is eaten up

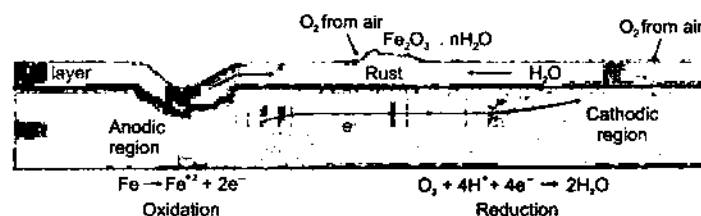



Fig. 7.6 Rusting of iron.

Q.14 Describe the methods for the prevention of corrosion.

Ans: Corrosion:

Correction is a slow and continuous eating away of metal by the surrounding medium. The most common example of corrosion is rusting of iron.

Methods for the prevention of corrosion:

i. Removal of stains

The regions of stains in an iron rod act as the site for corrosion. If the surface of iron is properly cleaned and stains are removed. It would prevent corrosion.

ii. Paints and greasing

Greasing, Polishing or painting of the surface can prevent the corrosion of iron. With development of technologies, modern paints contain a combination of chemicals called stabilizers that provide protection against the corrosion in addition to prevention against the weathering and other atmospheric effects.

iii. Alloying

Alloy is a homogeneous mixture of one metal with one or more other metals or non-metals. Alloying of iron with other metals has proved to be very successful technique against rusting. The best example of alloying is the 'stainless steel', which is a good combination of iron, chromium and nickel.

iv. Metallic coating

The best method for protection against the corrosion of metals exposed to acidic conditions is coating the metal. Corrosion resistant metals like Zn, Sn and Cr are coated on the surface of iron to protect it from corrosion. It is the most widely applied technique in the food industry where food is 'tin-packed'. The containers of iron are coated with tin or chromium to give it a longer life.

Method of metallic coating:

a. Physical Methods

b. Chemical method (Electroplating)

1. Zinc coating or Galvanizing

"The process of coating a thin layer of zinc on iron is called galvanizing."

This process is carried out by dipping a clean iron sheet in a zinc chloride bath and then heating it. After this iron sheet is removed, rolled into molten zinc metal bath and finally air-cooled. Advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

2. Tin Coating or Tinning:

The process of coating a thin layer of tin on iron is called tinning.

It involves the dipping of the clean sheet of iron in a bath of molten tin and then passing it through hot pairs of rollers. Such sheets are used in the beverage and food cans. Advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken. The tin protects the iron only as long as its protective layer remains intact. Once it is broken and the iron is exposed to the air and water, a galvanic cell is established and iron rusts rapidly.

Q.15 What is electroplating? Write down procedure of electroplating?

Ans: Electroplating:

Electroplating is depositing of one metal over the other by means of electrolysis.

Objective

This process is used to protect metals against corrosion and to improve their appearance, shine and beauty.

Principle:

Principle of electroplating is to establish an electrolytic cell in which anode is made of the metal to be deposited and cathode of the object on which metal is to deposit. The electrolyte is in aqueous solution of a salt of the respective metal.

Procedure for Electroplating:

i. Cleaning of metal:

In this process the object to be electroplated is cleaned with sand and washed with caustic soda solution.

ii. Anode: The anode is made of the metal, which is to be deposited like Cr, Ni. The

iii. Cathode: The cathode is made up of the object that is to be electroplated like some sheet made up of iron.

iv. Electrolyte: The electrolyte in this system is a salt of the metal being deposited. The

v. Electrolytic Tank: The electrolytic tank is made of cement, glass or wood in which anode and cathode are suspended.

The electrodes are connected with a battery. When the current is passed, the metal from anode dissolves in the solution and metallic ions migrate to the cathode and discharge or deposit on the cathode (object). As a result of this discharge, a thin layer of metal deposits on the object, which then pulled out and cleaned.

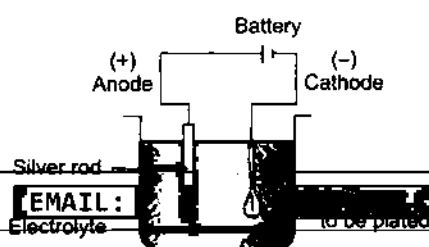



Fig. 7.7 Electroplating of an object.

Q.16 Describe electroplating of Silver, Chromium, Zinc and Tin in detail.

Ans: a. Electroplating of Silver

Principle:

The electroplating of silver is carried out by establishing an electrolytic cell.

Anode: The pure piece of silver strip acts as anode that is dipped in silver nitrate solution.

Cathode: The cathode is the metallic object to be coated such as silver spoon.

Chemical reaction: When the current is passed through the cell, the Ag^- ions dissolve at the anode and migrate towards the cathode where, they discharge and deposit on the object e.g. spoon. The chemical reaction can be represented as:

At anode: $Ag_{(s)} \longrightarrow Ag^+_{(aq)} + 1e^-$

At cathode: $Ag^+_{(aq)} + 1e^- \longrightarrow Ag_{(s)}$

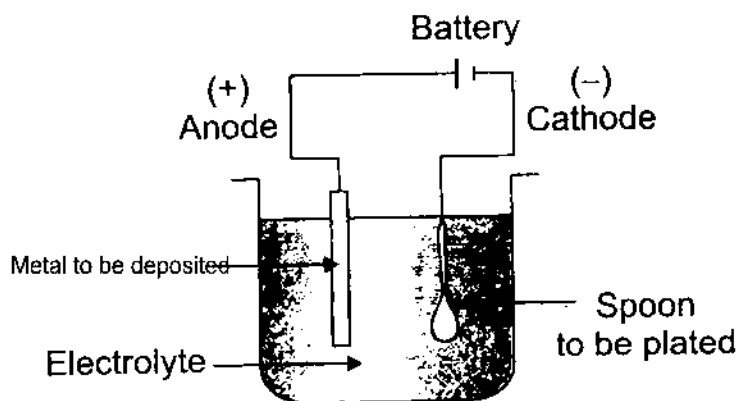
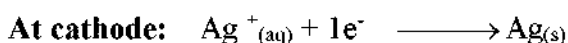
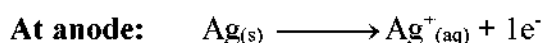



Fig. 7.7 Electroplating of an object.

Uses:

Common examples of silver plating are table wares, cutlery, jewelry and steel objects.

b. Electroplating of Chromium

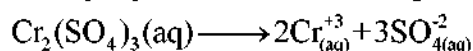
Principle: The electroplating of chromium is carried out by establishing an electrolytic cell.

Electrolyte: Aqueous solution of chromium sulphate containing a little sulphuric acid, acts as an electrolyte.

Cathode: The object to be electroplated acts as cathode

Anode: Anode is made of antimonial lead.

Chemical reactions: The electrolyte ionizes and provides Cr^{3+} ions, which reduce and deposit at cathode. Electrolyte produces the following ions.

$$Cr_2(SO_4)_3(aq) \longrightarrow 2Cr^{+3}_{(aq)} + 3SO^{-2}_{4(aq)}$$

Reactions at the electrodes are as follows.

At anode:

$$4OH^-_{(aq)} \longrightarrow 2H_2O_{(l)} + O_{2(g)} + 4e^-$$

At cathode:

$$Cr^{+3} + 3e^- \longrightarrow Cr$$

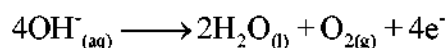
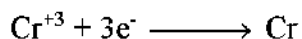
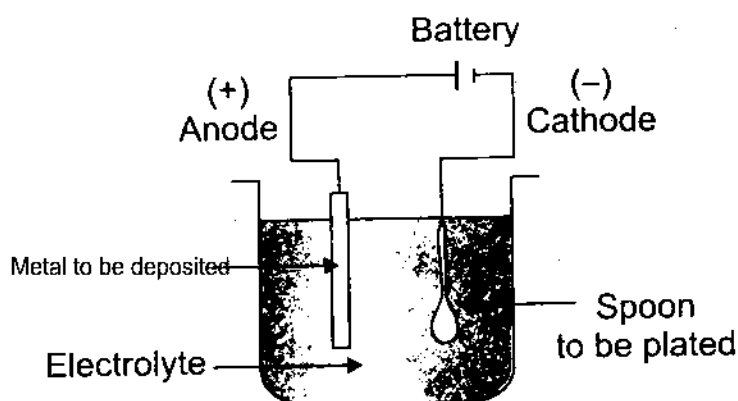



Fig. 7.7 Electroplating of an object.

Why steel is usually plated first with nickel or copper and then by chromium?

For practical convenience, the steel is usually plated first with nickel or copper and then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is stripped off. The nickel or copper provides adhesion and then chromium deposited over the adhesive layer of copper lasts longer. This type of electroplating resists corrosion and gives a bright silvery appearance to the object.

c. Electroplating of zinc

Principle:

The electroplating of zinc is carried out by establishing an electrolytic cell.

Cleaning of target metal: The target metal is cleaned in alkaline detergent type solutions, and it is treated with acid, in order to remove any rust or surface scales.

Anode: Zinc metal is used as anode.

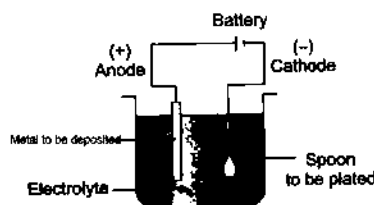



Fig. 7.7 Electroplating of an object.

Cathode: The metal to be plated is used as cathode.

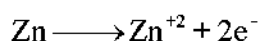
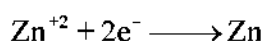
Electrolyte: Aqueous solution of zinc sulphate ($ZnSO_4$) is used as an electrolyte.

Next, the zinc is deposited on the metal by immersing it in a chemical bath containing electrolyte zinc sulphate. A current is applied, which results in zinc being deposited on the target metal i.e. cathode.

At anode:

$$Zn \longrightarrow Zn^{+2} + 2e^-$$

At cathode:

$$Zn^{+2} + 2e^- \longrightarrow Zn$$

d. Tin coating:

Principle:

The electroplating of tin is carried out by establishing an electrolytic cell.

Cleaning of target metal: The target metal is cleaned in alkaline detergent type solutions, and it is treated with acid, in order to remove any rust or surface scales.

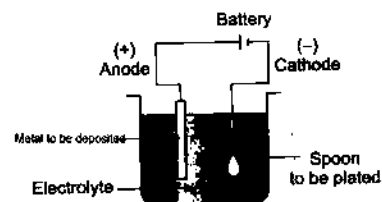



Fig. 7.7 Electroplating of an object.

Anode: Tin metal is used as anode.

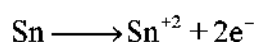
Cathode: The metal to be plated is used as cathode.

Electrolyte: Aqueous solution of tin sulphate ($SnSO_4$) is used as an electrolyte. Tin is usually electroplated on steel by placing the steel into a container containing a solution of tin salt. The teel is connected to an electrical circuit, acting as cathode. While the other electrode made of tin metal acts as anode. When an electrical current passes through the circuit, tin metal ions present in the solution deposit on steel.

At anode:

$$Sn \longrightarrow Sn^{+2} + 2e^-$$

At cathode:

$$Sn^{+2} + 2e^- \longrightarrow Sn$$

d. Electrolytic refining of copper

Impure copper is refined by the electrolytic method in the electrolytic cell.

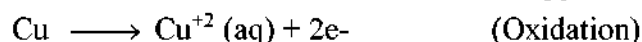
Anode: Impure copper acts as anode

Cathode: a pure copper plate acts as cathode

Electrolyte: Copper sulphate solution in water is used as an electrolyte.

Reaction at anode:

Oxidation reaction takes place at the anode. Copper atoms from the impure copper lose electrons to the anode and dissolve in solution as copper ions:

$$Cu \longrightarrow Cu^{+2}(aq) + 2e^- \qquad \text{(Oxidation)}$$

Reaction at cathode:

Reduction reaction takes place at the cathode. The copper ions present in the solution are attracted to the cathode. Where they gain electrons from the cathode and become neutral and deposit on the cathode.

$$Cu^{+2} + 2e^- \longrightarrow Cu \qquad \text{(Reduction)}$$

In the process impure copper is eaten up and purified copper atoms deposit on the cathode. 99.9% pure copper is obtained in this process.

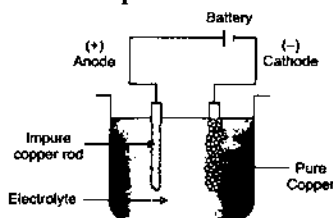



Fig 7.8 Refining of copper in an electrolytic cell.

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 7: Electrochemistry
Extra MCQ's

1. **Symbol of hydronium ion is:**
(a) H^+ (b) OH^- (c) H_3O^+ (d) none of these
2. **An example of non electrolyte is:**
(a) Glucose (b) aq NaCl (c) HCl (d) H_2SO_4
3. **An example of weak electrolyte is :**
(a) HNO_3 (b) HCl (c) H_2SO_4 (d) H_2CO_3
4. **During electrolysis _____ takes place at anode.**
(a) Catenation (b) oxidation (c) reduction (d) addition
5. **Which is not the characteristic of electrolyte?**
(a) Cheap (b) conductor (c) easily oxidized (d) soluble in water
6. **In dry cell ___acts as cathode:**
(a) Zn cup (b) graphite rod (c) paste (d) steel rod
7. **In Zn-Cu galvanic cell, Zn is dipped in:**
(a) $ZnSO_4$ (b) $Zn(NO_3)_2$ (c) $CuSO_4$ (d) both a & b
8. **In Zn-Cu galvanic cell, Zn is used as:**
(a) Cathode (b) anode (c) electrode (d) all of above
9. **In $Zn + Cu^{+2} \longrightarrow Zn^{+2} + Cu$, Zn is:**
(a) Oxidized (b) reduced (c) redoxed (d) decomposed
10. **Anions are ________ions.**
(a) Positive (b) neutral (c) negative (d) amphoeteric
11. **Cations are _______ ions.**
(a) Negative (b) positive (c) neutral (d) none of these
12. **Electrolysis of NaCl is done in the cell:**
(a) Electrolytic (b) voltaic (c) down's (d) faradays'
13. **Which one is not produced during electrolysis of aqueous NaCl?**
(a) Na^+ (b) OH^- (c) H_2O (d) Cl^-
14. **In pure water, out of ____only one molecule ionizes.**
(a) 6×10^6 (b) 6×10^8 (c) 6×10^{16} (d) 6×10^{12}
15. **Which one is weak electrolyte?**
(a) Citric acid (b) carbonic acid (c) tartaric acid (d) all of these
16. **Metallic, ammonium and hydrogen ions carry _____charges.**
(a) Negative (b) neutral (c) positive (d) partial positive
17. **Electrode through which electrons enter the external circuit:**
(a) Anode (b) Cathode (c) electrode (d) none of these

18. Electrode through which electrons leave the external circuit:
(a) Anode (b) Cathode (c) graphite (d) electrolyte

19. Which one is conductor?
(a) Naphthalene (b) paraffin wax (c) plastic (d) HCl

20. Which solution is not a conductor:
(a) NaCl (b) KI (c) $CO(NH_2)_2$ (d) $CuCl_2$

21. Rods through which electric current enters or leaves the cell:
(a) Protons (b) electrons (c) electrodes (d) all of these

22. Spontaneous redox reaction produce current in:
(a) Voltaic cell (b) electrolytic cell (c) galvanic cell (d) Both a & b

23. In an oxidation reaction electrons are:
(a) Absorbed (b) lost (c) moved (d) increased

24. In reduction reaction electrons are:
(a) Lost (b) absorbed (c) kept constant (d) all of these

25. Which of the following is a good electrolyte?
(a) NaCl (b) H_2SO_4 (c) NaOH (d) All of these

26. Which ionizes in small extent in water
(a) $Ca(OH)_2$ (b) NaCl (c) NaOH (d) H_2SO_4

27. Oxidation always takes place at
(a) Anode (b) Cathode (c) Both of them (d) Nome of them

28. Who invented first electrolytic cell?
(a) Berzelius (b) Volta (c) J. Dalton (d) Newton

29. In galvanic cell, cathode carries.
(a) Positive charge (b) Negative charge (c) No charge (d) Neutral charge

30. Cl_2 gas is formed, when Cl^- ions are
(a) Reduced (b) Oxidized (c) Removed (d) Reacted with metals

31. Which ion is not formed during is of aqueous sodium chloride?
(a) H^+ (b) H_2 (c) Cl_2 (d) NaOH

32. Which medium accelerates the process of rusting?
(a) Acidic (b) Basic (c) Buffer (d) Neutral

33. Stainless steel contains
(a) Nickel (b) Iron (c) Chromium (d) All of these

34. Which of the followings is a corrosion resistant metal?
(a) Fe (b) Zn (c) Sn (d) Sr

35. The electrolytic cell is made up of
(a) Cement (b) Glass (c) Wood (d) All of these

36. Which of the following is a common example of silver plating?
(a) Wares (b) Cutlery (c) Jewellery (d) All of these

37. Which metal has a great tendency to corrosion?
(a) Silver (b) Copper (c) Aluminum (d) All of these

38. Chemical formula of sodium thio-sulphate is

(a) Na_2SO_4 (b) Na_2SO_3 (c) $Na_2S_2O_3$ (d) $NaSO_4$

39. Sterling silver is an alloy of silver and
(a) Iron (b) Copper (c) Chromium (d) Aluminum

40. In HCl, oxidation number of H is
(a) –1 (b) +1 (c) +2 (d) –2

41. The oxidation number of all elements in free state is
(a) One (b) Two (c) Three (d) Zero

42. The oxidation number of Group-I elements is
(a) +1 (b) +2 (c) +3 (d) +4

43. The oxidation number of hydrogen in metal hydrides is
(a) +1 (b) –1 (c) +2 (d) –2

44. The oxidation number of oxygen is +2 in
(a) H_2O (b) OF_2 (c) Na_2O (d) HNO_3

45. The oxidation number of sulphur in H_2SO_4 is
(a) +1 (b) +4 (c) +6 (d) +8

ANSWER KEY

1	c	11	b	21	c	31	a	41	d
2	a	12	c	22	d	32	a	42	a
3	d	13	c	23	b	33	d	43	b
4	b	14	b	24	b	34	c	44	b
5	c	15	d	25	d	35	d	45	c
6	b	16	c	26	a	36	d		
7	a	17	a	27	a	37	c		
8	b	18	b	28	b	38	c		
9	a	19	d	29	a	39	b		
10	c	20	c	30	b	40	b		

Last Updated: November 2020
Report any mistake at freeilm786@gmail.com

Unit 8: Chemical Reactivity
Exercise Questions

Exercise Multiple Choice Question Answers:

1. **Metals can form ions carrying charges:**
(a) Uni-positive (b) Di-positive (c) Tri-positive (d) All of them
2. **Which one of the following metals burn with brick red flame when heated in air?**
(a) Sodium (b) Magnesium (c) Iron (d) Calcium
3. **Sodium is extremely reactive metal, but it does not react with:**
(a) Hydrogen (b) Nitrogen (c) Sulphur (d) Phosphorus
4. **Which one the following is the lightest and floats on water:**
(a) Calcium (b) Magnesium (c) Lithium (d) Sodium
5. **Pure alkali metals can be cut simply by knife but iron cannot because of alkali metals have:**
(a) Strong metallic bonding (b) Weak metallic bonding
(c) Non-metallic bonding (d) Moderate metallic bonding
6. **Which of the following is less malleable?**
(a) Sodium (b) Iron (c) Gold (d) Silver
7. **Metals lose their electrons easily because:**
(a) They are electronegative (b) They have electron affinity
(c) They are electropositive (d) Good conductors of heat
8. **Which one of the following is brittle?**
(a) Sodium (b) Aluminium (c) Selenium (d) Magnesium
9. **Which one of the following non-metal is lustrous?**
(a) Sulphur (b) Phosphorus (c) Iodine (d) Carbon
10. **Non-metalsare generally soft, but which one of the following is extremely hard?**
(a) Graphite (b) Phosphorus (c) Iodine (d) Diamond
11. **Which one of the following will not react with dilute HCl?**
(a) Sodium (b) Potassium (c) Calcium (d) Carbon

ANSWR KEY

1	d	3	b	5	b	7	c	9	c	11	d
2	d	4	d	6	b	8	c	10	d		

Exercise Short Question Answers

Q.1 Why reactivity of metals increases down the group?

Ans: Reactivity of metals depends upon its tendency to lose electrons which is dependent upon the size of atoms. Thus reactivity of metals increases down the group because of increasing atomic size and decreasing ionization energy.

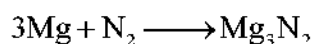
Q.2 Why reactivity of metals increases down the group?

Ans: Physical properties of metals:

i. Almost all metals are solids (except mercury).
ii. They have high melting and boiling points.
iii. They possess metallic luster.
iv. They are malleable and ductile.
v. They are good conductors of heat and electricity.
vi. They have high densities.

Q.3 Why nitrogen forms compounds with alkaline earth metals directly?

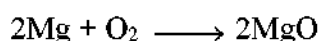
Ans: Nitrogen forms compounds with alkaline earth metals directly because alkaline earth metals from di-positive cations (M^{++}). They have high charge density and polarization power. They can polarize nitrogen atoms easily and produce stable covalent nitrides with nitrogen.

$$3Mg + N_2 \longrightarrow Mg_3N_2$$

Q.4 Why the second ionization energy of magnesium is higher than the first one?

Ans: Second Ionization energy of magnesium is higher than the first one because after the removal of its electron nuclear charge increases and atomic size decreases. The remaining electrons will be attracted by the nucleus more strongly.

Q.5 How oxygen reacts with group II A metals?

Ans: They are less reactive towards oxygen and they form oxides on heating.

$$2Mg + O_2 \longrightarrow 2MgO$$

Q.6 What is relationship between electro positivity and ionization energy

Ans: Electropositivity depends upon the ionization energy which in turn depends upon size and nuclear charge of the atom. Small sized atoms with high nuclear charge have high ionization energy. Atoms having high ionization energy are less electropositive or metallic.

Q.7 Why electro positivity decreases from left to right in a period?

Ans: Electropositivity decreases across the period from left to right in the periodic table because. Size of atoms decreases due to increase in nuclear charges.

Q.8 How electro positivity depends upon size and nuclear charge of an atom?

Ans: Electropositivity depends upon size and nuclear charge of an atom because when the size of atoms increases, electropositivity increases as it becomes easier to lose electrons. It also depends upon nuclear charge. If nuclear charge increases the electropositivity decreases because it becomes difficult to remove the electrons from outermost shell.

Q.9 Why ionization energies of alkaline earth metals are higher than alkali metals?

Ans: Ionization energies of alkalme earth metals are higher than alkalimetals because the atomic size of alkalme earth metals one smaller and greater nuclear charge.

Q.10 Why silver and gold are least reactive?

Ans: Silver and gold are least reactive because these metals do not lose their electrons easily. They do not have the tendency to make cations.

Q.11 Can pure gold be used for making ornaments? If not why?

Ans: No, pure gold cannot be used for making ornaments because gold is too soft to be used as such. It is always alloyed with copper, silver or some other metal.

Q.12 Why copper is used for making electrical wires?

Ans: Copper is used for making of electrical wires because it is a good conductor of electricity and can easily be drawn into wires.

Q.13 What is the trend of variation in densities of alkali metals?

Ans: Densities of alkali metals increase down the group in the periodic table due to increase atomic mass.

Q.14 Which metal is used for metal work?

Ans: Metal work means objects that are made in an artistic and skilful way. Copper metal is used in metal work because it is easily workable. It is used in many ornaments, plumbing, roafing and other operations.

Q.15 Why magnesium is harder than sodium?

Ans: Magnesium is harder than sodium because in magnesium metallic bonding is stronger, than sodium. Magnesium involves 2 valence electrons in metallic bonding as compared to sodium which involves only one valence electron. Moreover magnesium has smaller atomic size and high ionization energy.

Q.16 Why calcium is more electropositive than to magnesium?

Ans: Calcium is more electropositive than magnesium because calcium has larger size and greater nuclear charge than magnesium and in turn lower ionization energy than magnesium.

Q.17 Why ionization energy of Na is less than Mg?

Ans: Ionization energy of Na is less than Mg because Na has larger size and lower nuclear charge than Mg. It results in less nuclear attraction on valance electron in sodium.

Q.18 Why the ionization energy of Na is more than K?

Ans: The ionization energy of sodium is more than K because down the group electropositive character increases and ionization energy decreases. It becomes easier for potassium to

lose electrons than sodium.

Exercise Long Question Answers

Q.1 Compare and contrast the properties of alkali and alkaline earth metals.

Ans: See Q. No. 5 (Subjective Part, Long Questions Answers)

Q.2 Discuss the inert character of silver and gold.

Ans: Silver and gold are inert metals because they both are very less electropositive and do not lose electrons easily.

Inert Character of Silver:

Silver is a white lustrous metal. Formation of thin layer of oxide or sulphide on its surface makes it relatively un reactive. Under normal conditions of atmosphere, air does not affect silver. It tarnishes in presence of sulphur containing compounds like H_2S.

Inert Character of Gold:

Gold is a yellow soft metal. Gold is very non reactive or inert metal. It is not affected by atmosphere. It is not even affected by any single mineral acid or base. It dissolves only in Aqua Regia. Because of its inertness in atmospheres it is an ornamental metal as well as used in making coins.

Q.3 Why cations are smaller and anions are bigger in size than their respective neutral atoms.

Ans: **Small Size of Cation (Positive Ion) than its neutral atom:**

Cations are smaller than their corresponding neutral atoms because of two reasons.

i) The removal of one or more electrons from a neutral atom usually, results in the loss of the outer most shell.

ii) The removal of electrons causes an imbalance in proton-electron ration thus a cation has smaller number of electrons than its parent atom with the decrease in number of electrons the magnitude of effective nuclear charge increases, which pulls the electrons cloud of the cation near to the nucleus and thus makes the cation smaller in size than its parent neutral atom.

Examples:

The radius of Na is 186pm whereas ionic radius of cation (Na^+) is 102pm.

Large size of anion (Negative Ion) than its neutral atom

A negative ion is always bigger than its parent atom the reason is that the addition of one or more electrons in the shell of a neutral atom enhance the repulsion between the electron causing the expansion of the shell. The added electrons reduce the attraction of nucleus to the electron that is why the size of anion increases as compared to the neutral atom.

Examples:

Atomic size of Flourine (F) is 71pm whereas anionic size of Fluorine (F^-) is 136 pm.

Q.4 Discuss why hardness and softness of a metal depends upon its metallic bonding.

Ans: The softness and hardness of a metal depends upon the metallic bonding. The strength of the metallic bonds upon the number of valence electrons that each atom contributes for the metallic bonding.

Hardness of a Metal:

Some metals have strong metallic bond due to the greater number of valence electrons in the metal atoms. Such metals are Cord.

Examples:

Magnesium metal has strong metallic bond as compared to sodium metal therefore magnesium is harder than sodium metal.

Softness of a metal

Some metals have weak metallic bond due to the less number valance electrons in the metal atoms. Such metals are soft.

Examples:

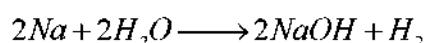
Sodium has weak metallic bond as compared to magnesium metal that's why it is soft as compared to magnesium. It has low melting point and can easily be cut with knife.

Q.5 Give the reaction of sodium with; H_20, O_2, Cl_2 and H_2

Ans:

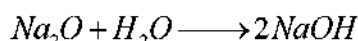
i) Reaction of Sodium with H_2O.

Sodium reacts with water vigorously at room temperature to give strong alkaline solution and hydrogen gas..

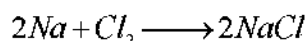
$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

ii) Reaction f sodium with O_2:

Sodium immediately tarnishes in air giving sodium oxide which forms strong alkali in water.

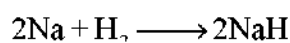
$$4Na + O_2 \longrightarrow 2Na_2O$$

$$Na_2O + H_2O \longrightarrow 2NaOH$$

iii) Reaction of sodium with Cl_2:

Sodium reacts violently with chlorine at room temperature to give sodium chloride.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

iv) Reaction of sodium with H_2:

Sodium reacts with hydrogen, at high temperature to form sodium hydride.

$$2Na + H_2 \longrightarrow 2NaH$$

Q.6 What are physical properties of calcium metal? Give its uses.

Ans: Physical properties of calcium metal

Following are physical properties of calcium.

1. Calcium is silvery grey and fairy harder.
2. Its density is 1.55g cm^{-3}
3. It is malleable and ductile.
4. It is good conductor of heat and electricity.
5. Its melting point is 851°C and boiling point is 1484°C.
6. Its flame colour is brick red.
7. Its first ionization energy is 590kj mol^{-1} and second ionization energy is 1145kgmol^{-1}.

Uses of calcium

1. It is used to remove sulphur from petroleum products.
2. it is used as reducing agent to produce Cr, u and Zr,

Q.7 Write down the chemical properties of the non-metals?

Ans: See Q. No. **10** (Subjective Part, Long Questions Answers)

Q.8 Compare the physical properties of metals and non-metals

Ans:

Metals	Non Metals
1. All metals are solids except mercury.	1. Non metals are solid, liquid and gases.
2. They have high melting and boiling point.	2. They have low melting and boiling point.
3. They have metallic luster and can be polished.	3. They do not have metallic luster and cannot be polished they have dull surface.
4. They are malleable and ductile.	4. They are not malleable and ductile.
5. They are good conductor of heat and electricity.	5. They poor conductors of heat and electricity.
6. They have high densities.	6. They have low densities.
7. They are usually hard.	7. They are usually soft.

Q.9 How you can compare the softness and hardness of metals?

Ans: Softness and hardness of metals depends upon the strength of metallic bond present in them.

Dependency of Metallic bond

The strength of a metallic bond depends upon the following factors.

i. Charge present on positive metallic ion.

ii. Number of mobile electrons set free by each atom.

Softness of metals:

Metals having weak metallic bond are soft metals, such metals have low melting points,

boiling points, densities etc.

Example:

Sodium is a soft metal due to weak metallic bond. It can be cut with a knife. Its melting point is very low as compared to other metals.

Harness of metals:

Metals having strong metallic bond are hard metals. Such metals have high melting points, boiling points densities etc.

Example:

Magnesium is a hard metal due to strong metallic bond, its melting point is 650°C which is very high as compared to sodium.

Q.10 Give the chemical properties of magnesium and its uses

Ans: Chemical properties of magnesium:

i. Reaction with water

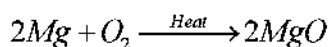
Magnesium reacts with water less rigorously and on heating produces weak base.

$$Mg + H_2O \longrightarrow MgO + H_2$$

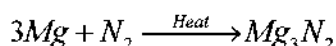
$$MgO + H_2O \longrightarrow Mg(OH)_2$$

ii. Reaction with oxygen

Magnesium reacts with oxygen on heating and magnesium oxide is formed.

$$2Mg + O_2 \xrightarrow{Heat} 2MgO$$

iii. Reaction with Nitrogen

Magnesium form stable nitride when heat end with nitrogen.

$$3Mg + N_2 \xrightarrow{Heat} Mg_3N_2$$

Uses of Magnesium:

i. Magnesium used in flash light bulbs and in fire works.

ii. It is used in the manufacture of light alloys.

iii. Magnesium ribbon is used in thermite process to ignite aluminium powder.

iv. Magnesium is used as anode for prevention of corrosion. .

Q.11 Write a comprehensive note on the electropositive character of metals?

Ans: See Q. No. **2** (Subjective Part, Long Questions Answers)

Q.12 Compare the ionization energies of alkali and alkaline earth metals.

Ans: See Q. No. **3** (Subjective Part, Long Questions Answers)

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 8: Chemical Reactivity
Long Questions

8.1 Metals

Q.1 What are metals? How are they categorized? Write down their physical and chemical properties.

Ans: Metals:

"Metals are the elements (except hydrogen) which are electropositive and form cations by losing electrons."

Examples: Sodium, Potassium, iron, Silver etc.

Importance of metals:

Things like aeroplanes, trains, building frames, automobiles or even different machines and tools are due to different properties of metals.

Position of metals in periodic table:

They are present on the lower and left side of the periodic table.

Fig. 8.1 Some common metals and non-metals.

Categories of metals:

Metals can be categorized as:

i. **Very reactive:** potassium, sodium, calcium, magnesium and aluminum.
ii. **Moderately reactive:** zinc, iron, tin and lead.
iii. **Least reactive or noble:** copper, mercury, silver and gold.

Physical characteristics:

i. Almost all metals are solids (except mercury)
ii. They have high melting and boiling points.
iii. They possess metallic luster and can be polished.
iv. They are malleable (can be hammered into sheets), ductile (can be drawn into wires) and give off a tone when hit.
v. They are good conductor of heat and electricity.
vi. They have high density.
vii. They are hard (except sodium and potassium)

Chemical properties:

i. They easily lose electrons and form positive ions.
ii. They readily react with oxygen to form basic oxides.
iii. They usually form ionic compounds with non-metals.
iv. They have metallic bonding.

DO YOU KNOW

- The most abundant metal is aluminum
- The most precious metal is platinum
- The most useable metal is iron
- The most reactive metal is cesium
- The most valuable metal is uranium
- The lightest metal is lithium (d = 0.53 g cm^{-3})
- The heaviest metal is osmium (d = 22.5 g cm^{-3})
- The least conductor of heat is lead.
- The best conductor metals are silver and gold
- The most ductile and malleable metals are gold and silver

Q.2 Write a compressive note on the electropositive character of metals

Ans: Electropositive character / Metallic character:

"Metals have the tendency to lose their valance electrons. This property of a metal is termed as electropositivity or metallic character or electropositive character."

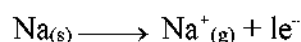
Valency of metals:

"The number of electrons lost by an atom of a metal is called its valency."

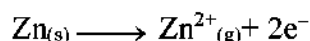
The more easily a metal loses its electrons is the more electropositive.

Example:

i. Sodium (Na) atom can lose 1 electron to form a positive ion

$$Na_{(s)} \longrightarrow Na^{+}_{(g)} + 1e^{-}$$

So the valency of sodium metal is1.

ii. Zinc (Zn) metal can lose 2 electrons from its valence shell. Therefore, its valency is 2.

$$Zn_{(s)} \longrightarrow Zn^{2+}_{(g)} + 2e^{-}$$

Trends of Electropositivity:

a. Trends in Groups:

Electropositive character increases down the group because size of atoms increases.

Example:

Lithium metal is less electropositive than sodium which is in turn less electropositive than potassium.

b. Trends in Periods:

Electropositive character decreases across the period from left to right in periodic table because size of atoms decrease due to increase of nuclear charge.

It means elements in the start of a period are more metallic. This character decreases as we move from left to right along the period.

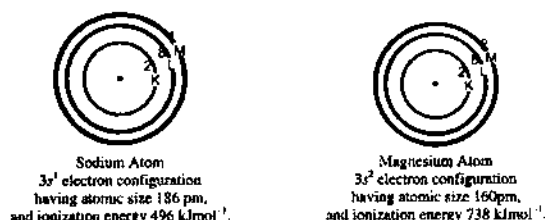
Q.3 (Ex. Q.12) compare the ionization energies of alkaline earth metals are what is relationship between electropositive and ionization energy.

Ans: Electropositive and ionization energy.

Dependence of electropositive character:

Electropositive character depends upon the ionization energy which in turn depends on size and nuclear charge of the atom. Small sized atoms with high nuclear charge have high ionization energy. In this way atoms having high ionization energy are less electropositive or metallic. That is the reason alkali metals have the largest size and the lowest ionization energy in their respective periods. Therefore, they have the highest metallic character.

Sodium Atom
$3s^1$ electron configuration
having atomic size 186 pm,
and ionization energy 496 kJmol^{-1}.

Magnesium Atom
$3s^2$ electron configuration
having atomic size 160pm,
and ionization energy 738 kJmol^{-1}.

Example: Comparison of sodium and magnesium metals

sodium	magnesium
• Sodium Atom $3s^1$ electron configuration having atomic size 186 pm	• Magnesium Atom $3s^2$ electron configuration having atomic size 160pm
• Ionization energy of sodium is 496 kjmol^{-1}.	• Ionization energy of magnesium 1450 kjmol^{-1}.

Comparison of second ionization energy and first ionization energy of alkaline earth metals:

Ionization energy of magnesium is high but the 2nd ionization energy of magnesium is very high. It becomes very difficult to remove second electron from the Mg^{+} ion as nuclear charge attracts the remaining electrons strongly. As a result of this attraction size of the ion decreases.

Similarly all the elements of alkaline earth metals have high ionization energies as compared to alkali metals

Metal	Atomic Number	Electronic Configuration	IE	Metal	Atomic Number	Electronic Configuration	IEI	IE2
Li	3	[He] 2 s^1	520	Be	4	[He] $2s^2$	899	1787
Na	11	[Ne] 3 s^1	496	Mg	12	[Ne] $3s^2$	738	1450
K	19	[Ar] 4 s^1	419	Ca	20	[Ar] $4s^2$	590	1145
Rb	37	[Kr] 5 s^1	403	Sr	38	[Kr] $5s^2$	549	1064
Cs	55	[Xe] 6 s^1	376	Ba	56	[Xe] $6s^2$	503	965

Note: *Low ionization energies of alkali metals make them more reactive than alkaline earth metals.*

Q.4 How you can compare physical properties of alkali and alkaline earth metals?

Ans: Comparison of physical properties of alkali and alkaline earth metals:

Property	Sodium	Magnesium	Calcium
Appearance	Silvery white having a metallic luster, very soft and can be cut with knife	Silvery white and hard	Silvery grey and fairly harder
Atomic size, ionic size (pm)	186, 102	160, 72	197,99
Relative density	0.98 g cm^{-3} Floats on water	1.74 g cm^{-3}	1.55 g cm^{-3}
Malleability	very malleable and ductile	Malleable and ductile	Malleable and ductile
Conductivity	Good conductor of heat and electricity	Good conductor of heat and electricity	Good conductor of heat and electricity
M.P	97°C	650°C	851°C
B.P	883°C	1090 °C.	1484°C
Ionization energy	496 kJ/mol	738, 1450 kJ/mol	590,1145 kJ/mol
Flame in air	Golden yellow	Brilliant white	Brick red

Q.5 Describe reactivates of alkali and alkaline earth metals,

OR

Compare chemical properties and reactivates of akali and alkaline earth metals?

Ans: Alkali metals:

"The elements in Group 1 (Li, Na, K, Rb, Cs, Fr) of the periodic table are called 'Alkali' metals".

Properties:

i. Alkali metals are extremely reactive elements because of their ns^1 valence shell electronic configuration.

ii. There is only one electron in their valence shell, it can be easily given out.

iii. They are always found in nature as cations with + I oxidation state.

iv. They readily form salts with non-metals.

Alkaline earth metals

"The elements in group 2 (Be, Mg, Ca, Sr, Ba, Ra) are called alkaline earth metals."

Properties:

i. The alkaline earth metal atoms are smaller and have more nuclear charge.

ii. They have two electrons in their valence shells.

iii. They are also reactive elements because of ns^2 valence shell electronic

configuration but less reactive than alkali metals because of small size and more nuclear charge.

Comparison of chemical properties and reactivates:

Alkali Metals	Alkaline Earth Metals
1. Occurrence They are very reactive and always occur in combined form	They are fairly reactive and also occur in combined form
2. Electropositivity These are highly electropositive. They have ionization energy values ranging from 520 $kJmol^{-1}$ for Li to 376 $kJmol^{-1}$ for Cs.	They are less electropositive. They have ionization energy values ranging from 1757 $kJmol^{-1}$ for Be to 965 $kJmol^{-1}$ for Ba.
3. Reaction with water They react with water vigorously at room temperature to give strong alkaline solution and hydrogen gas $2Na+2H_2O \longrightarrow 2Na\,OH+H_2$	They react with water less vigorously and on heating they produce weak bases $Mg + H_2O \longrightarrow MgO + H_2$ $MgO + H_2O \longrightarrow Mg\,(OH)_2$
4. Reaction with Oxygen. They immediately tarnish in air giving their oxides which form strong alkalies in water $4Na + O_2 \longrightarrow 2Na_2O$ $Na_20+H_2O \longrightarrow 2NaOH$	They are less reactive towards oxygen and oxides are formed on heating $2Mg+O_2 \longrightarrow 2MgO$
5. Reaction with Hydrogen They form ionic hydrides with H_2 at high temperature $2M+H_2 \longrightarrow 2MH$	They give hydrides under strong conditions of temperature and pressure $Ca + H_2 \longrightarrow CaH_2$
6. Reaction with Halogens They react violently with halogens at room temperature to give halides $2Na+Cl_2 \longrightarrow 2NaCI$	They react slowly with halogen to give their halides $Ca + Cl_2 \longrightarrow CaCl_2$
7. Reaction with Nitrogen They do not form nitrides directly	They form stable nitrides when heated with nitrogen $3Mg+N_2 \longrightarrow Mg_3N_2$
8. Reaction with Carbon They do not react with carbon directly	They give stable carbide on heating with carbon. $Ca+2C \longrightarrow CaC_2$

Q.6 Write down the uses of sodium, magnesium and calcium.

Ans: Uses of sodium:

i. Sodium-potassium alloy is used as a coolant in nuclear reactors.

ii. It is used to produce yellow light in sodium vapour lamps.

iii. It is used as a reducing agent in the extraction of metal like Ti.

Uses of magnesium:

i. Magnesium is used in flash lights and in fireworks.
ii. It is used in the manufacture of light alloys.
iii. Magnesium ribbon is used in Thermite process to ignite aluminium powder
iv. Magnesium is used as anode for prevention of corrosion.

Uses of calcium:

i. It is used to remove sulphur from petroleum products.
ii. It is used as reducing agent to produce Cr, U and Zn.

Q.7 Explain transition metals and inertness of Noble metals.

Ans: Transitions metals:

The elements in which d or f-orbital are in the process of filling, constitute a group of metals called transition metals.

Types of transition elements:

i. Outer transitions elements **ii. Inner transition elements**

i. Outer transitions elements:

The elements in which d-orbital are in the process of filling, constitute a group of metals called outer transition metals or d-group elements.

They exhibit a variety of oxidation states 'transition metals' of 4^{th}, $5^{th}$and the 6^{th} period of the periodic table. There are three series of transition elements; each series consisting of ten element

ii. Inner transition elements:

The elements in which f-orbital are in the process of filling, constitute a group of metals called inner transition metals or d-group elements.

There are 2 series of inner transitions elements placed at the bottom of the periodic table these series are lanthanides and actinides.

Inertness of Nobel Metals:

Chemical behavior of the first transition series is similar to active metals except copper. Three transition metals belonging to group 11 are copper, silver and gold.

Gold and silver are relatively inactive metals because they do not lose electrons easily and are called nobel metals.

Fig. 8.2 The Transition Elements in the Periodic Table.

Q.8 Write a note on properties and uses of silver, gold and platinum.

Ans: a. Properties of Silver:

i. It is white lustrous metal.

ii. It is an excellent conductor of heat and electricity.

iii. It is also highly ductile and malleable metal.

iv. Its polished surfaces are good reflectors of light.

v. Formation of thin layer of oxide or sulphide on its surface makes it relatively unreactive.

vi. Under normal conditions of atmosphere, air does not affect silver.

vii. It tarnishes in presence of sulphur containing compounds like H_2S.

Uses of silver:

i. Being very soft metal, it is rarely used as such.

ii. Alloys of silver with copper are widely used in making coins, silver-ware and ornaments.

iii. Compounds of silver are widely used in photographic films and dental preparations.

iv. Silver also has important applications in mirror industry.

b. Properties of Gold:

i. It is a yellow soft metal.

ii. It is most malleable and ductile of all the metals.

iii. One gram of gold can be drawn into a wire of one and a half kilo meter long.

iv. Gold is very non-reactive or inert metal.

v. It is not affected by atmosphere.

vi. It is even not affected by any single mineral acid or base.

Uses of Gold:

i. Because of its inertness in atmosphere, it is an ornamental metal as well as used in making coins.

ii. Gold is too soft to be used as such.

iii. It is always alloyed with copper, silver or some other metal.

Composition of pure gold

Purity of gold is shown by carats.

The number of parts by weight of gold that is present in 24 parts of alloy is called carats.

Twenty four carat gold is pure. 22 carats **gold** means that 22 parts pure gold is alloyed with 2 parts of either silver or copper for making ornaments and jewelry.

White gold

White gold is an alloy with Palladium nickel or zinc.

c. Properties and Uses of Platinum:

i. It is used to make jewelry items because of its unique characteristics like colour, beauty, strength, flexibility and resistance to tarnish.

ii. It provides a secure setting for diamonds and other gemstones, enhancing their brilliance.

iii. Platinum alloyed with palladium and rhodium are used as catalyst in auto- mobiles as catalytic convertor.

iv. They convert most of the gases being emitted by vehicles into less harmful carbon dioxide, nitrogen and water vapour. .

v. Platinum is used in the production of hard disk drive coatings and fibre optic cables.

vi. Platinum is used in the manufacturing of fibre glass reinforced plastic and glass for liquid crystal displays (LCD).

Q.9 What are non metals? Explain electronegative characteristics of non-metals

Ans: Non-Metals:

"The elements which form negative ions (anions) by gaining electrons are called non-metals"

Non-metals are electronegative in nature and form acidic oxides.

Examples: Oxygen, Sulphur, Phosphorous and Nitrogen etc.

Non- metallic character:

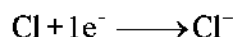
"The tendency of an element to gain electrons and from negative ions is called non-metallic character or electronegative character or electronegativity."

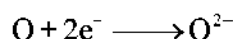
Valency of non-metals:

The valency of some non-metals depends upon the number of electrons accepted by them.

Examples:

i. Valency of chlorine atom is 1, as it accepts only I electron in its outermost shell.

$$Cl + 1e^- \longrightarrow Cl^-$$

ii. Oxygen atom can accept 2 electrons, therefore, its valency is 2.

$$O + 2e^- \longrightarrow O^{2-}$$

Dependence of non-metallic character:

The non-metallic character depends upon the electron affinity and electro negativity of the atom. Small size elements having high nuclear charge are electronegative in nature. They have high electron affinity. Therefore, they possess non-metallic nature.

Trend of non-metallic character:

a. Trends in groups:

Non-metallic character decreases in a group downward.

b. trends in periods:

It increases in a period from left to right up to halogens. That is the reason fluorine is the most non-metallic in character.

Example of non-metals:

The non-metals are, therefore, elements in Group-14(Carbon), Group-15 (nitrogen and phosphorus), Group-I6 (oxygen, sulphur and selenium) and in group-17 halogens (fluorine, chorine, bromine and iodine) of the periodic table.

Fig. 8.3 The Non-Metals in Periodic Table

Q.10 What are physical and chemical properties of non metals?

Ans: Physical properties of non-metals:

Physical properties of non-metals change gradually but uniquely in a group of non- metals. Non-metals usually exist in all three physical states of matter. The non-metals at the top of the group are usually gases while others are either liquids or solids.

Major physical properties of non-metals:

i. Solids non-metals are brittle (break easily).

ii. Non-metals are non-conductor of heat and electricity (except graphite).

iii. They are not shiny, they are dull except iodine (it is lustrous like metals).

iv. They are generally soft (except diamond).

v. They have low melting and boiling points (except silicon, graphite and diamond).

vi. They have low densities.

Chemical properties of non-metals:

i. Their valence shells are deficient of electrons, therefore, they readily accept electrons to complete their valence shells and become stable.

ii. They form ionic compounds with metals and covalent compounds by reacting with other non-metals e.g. CO_2, NO_2, etc.

iii. Non-metals usually do not react with water.

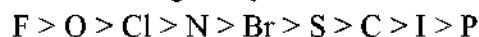
iv. They do not react with dilute acids because non-metals are itself electron acceptors.

v. They form acidic oxides.

Trend of electro negativity:

Electro negativity of first member of group 14, 15, 16 and 17 are higher than that of other members of the group decreasing their electro negativity.

The decreasing order of electronegativity is as under.

$$F > O > Cl > N > Br > S > C > I > P$$

Q.11 What are halogens? Compare the reactivity of the halogens in detail?

Ans: Halogens:

"Elements of Group-17 of the periodic table consist of fluorine, chlorine, bromine, iodine and astatine. They are collectively called halogens."

Physical states of halogens:

Fluorine and chlorine exist as diatomic gases at room temperature. The intermolecular forces of attraction increase downward in the group due to the increase in the size of atom. Bromine exists as a liquid and iodine as solid.

Physical properties of halogens, physical characteristics of halogens

Element	Atomic No.	Electronic Configuration	Colour	Melting Point (K)	Boiling Point (K)	Electro negativities
F	9	$[He]2s^2p^5$	Pale Yellow	53	85	4.0

Cl	17	$[Ne]3s^3p^5$	Greenish	172	238	3.2
Br	35	$[Ar]4s^2 4p^5$	Reddish Brown	266	332	3.0
I	53	$[Kr]5s^25p^5$	Purple Black	387	457	2.7

General characteristics:

- **i.** In general their valence shell electronic configuration is ns^2np^5.
- **ii.** Halogens have only one electron deficient in their valence shell; either they can readily accept an electron from a metal or they can share an electron with other non-metals.
- **iii.** Halogens form ionic bonds with metals and covalent bond with non-metals.

Q.12 Give the chemical properties of halogens.

OR

Describe important reactions of halogens.

Ans: Reactions of Halogens:

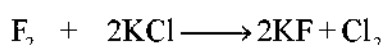
i. Oxidizing properties of halogens:

"A substance that oxidizes another substance by taking its electrons is called an oxidizing agent."

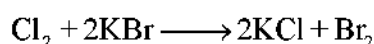
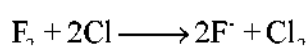
All halogens are oxidizing agent. Fluorine is the strongest oxidizing element while iodine is the least i.e. Cl is mild oxidizing agent. Fluorine will oxidize any of halide ion (X^{-1}) in solution and changes itself to F^- ion. Chlorine will displace Br^- and I^-ion from their salt solutions and oxidize them to bromine and iodine.

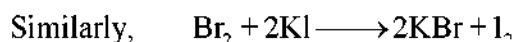
Order of reactivity of halogens is as follows

$$F > Cl > Br > I$$

$$F_2 + 2KCl \longrightarrow 2KF + Cl_2$$

More reactive Less reactive

$$F_2 + 2Cl \longrightarrow 2F^- + Cl_2$$

$$Cl_2 + 2KBr \longrightarrow 2KCl + Br_2$$

Solution turns from colourless to reddish brown

Similarly, $Br_2 + 2KI \longrightarrow 2KBr + I_2$

Actually a more reactive halogen can displace and oxidize are less reactive halogen.

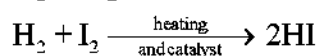
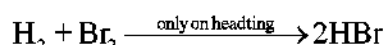
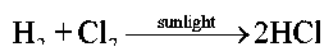
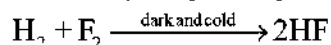
ii. Reaction with hydrogen:

All halogens (X_2) combine with hydrogen to give hydrogen halides (HX).

$$H_2 + F_2 \xrightarrow{\text{dark and cold}} 2HF$$

$$H_2 + Cl_2 \xrightarrow{\text{sunlight}} 2HCl$$

$$H_2 + Br_2 \xrightarrow{\text{only on headting}} 2HBr$$

$$H_2 + I_2 \xrightarrow[\text{and catalyst}]{\text{heating}} 2HI$$

Trend of chemical reactivity of halogen:

- **i.** The chemical affinity for H_2 decreases down the group from F_2 to Br_2
- **ii.** Fluorine combines with hydrogen even in the dark and cold state. Chlorine reacts with

hydrogen in the presence of sunlight.

iii. Bromine and iodine react with hydrogen only on heating.

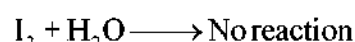
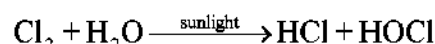
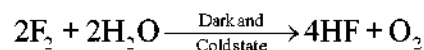
iii. Reaction with water:

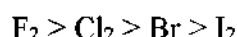
Flourine (F_2) decomposes water in cold state and in dark. Chlorine decomposes water in presence of sunlight. Bromine only react with water under special conditions. Iodine does not give this reaction.

$$2F_2 + 2H_2O \xrightarrow[\text{Cold state}]{\text{Dark and}} 4HF + O_2$$

$$Cl_2 + H_2O \xrightarrow{\text{sunlight}} HCl + HOCl$$

$$Br_2 + H_2O \xrightarrow{\text{heat}} HBr + BOBr$$

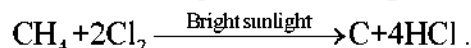
$$I_2 + H_2O \longrightarrow \text{No reaction}$$

iv. Reaction with methane:

$$F_2 > Cl_2 > Br > I_2$$

Reaction with fluorine: (F_2) reacts violently with methane (CH_4) in dark.

Reaction with chlorine:

a. In dark: Chlorine (Cl_2) does not react with methane in dark.

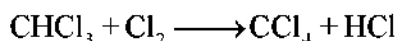
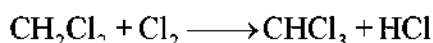
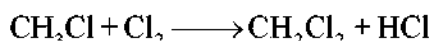
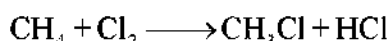
b. In bright sunlight: However the presence of bright sunlight the reaction is violent.

$$CH_4 + 2Cl_2 \xrightarrow{\text{Bright sunlight}} C + 4HCl.$$

c. In diffused sunlight: In presence of diffused sunlight the reaction of chlorine with methane is slow and gives series of compounds e.g. CH_3Cl, CH_2Cl_2, $CHCl_3$ and CCl_4.

$$CH_4 + Cl_2 \longrightarrow CH_3Cl + HCl$$

$$CH_3Cl + Cl_2 \longrightarrow CH_2Cl_2 + HCl$$

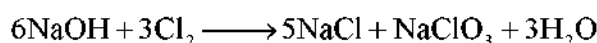
$$CH_2Cl_2 + Cl_2 \longrightarrow CHCl_3 + HCl$$

$$CHCl_3 + Cl_2 \longrightarrow CCl_4 + HCl$$

v. Reaction with Sodium hydroxide:

Chlorine reacts with **cold dilute NaOH** to give sodium hypochlorite

$$2NaOH + Cl_2 \longrightarrow NaCl + NaOCl + H_2O$$

Cl_2 reacts with **hot concentrated NaOH** to give sodium chloride in

$$6NaOH + 3Cl_2 \longrightarrow 5NaCl + NaClO_3 + 3H_2O$$

Q.13 What is the significance of non metals in daily life?

Significance of Non-metals:

Although non-metals are fewer than metals, yet they are highly significant. They are equally important for human beings, animals and plants. In fact, life would not have been possible without the presence of non-metals on earth.

i. As major components of earth's crust oceans and atmosphere:

Major components of earth's crust, oceans and atmosphere are non-metals: oxygen has the highest percentage in earth' crust (47%) and oceans (86%) and it is second (21%) to nitrogen in atmosphere. It indicates the importance of oxygen in nature.

Maintenance of balance of non-metals:

To maintain the balance for the amount of non-metals in nature, different cycles like water cycle, nitrogen cycle etc have been established naturally.

ii. As essential component of body:

Non-metals are essential part of the body structure of all living things.

Examples:

a. Human body is made up of about 28 elements. But about 96% of the mass of the human body is made up of just 4 elements i.e. oxygen 65%, carbon 18%, hydrogen 10% and nitrogen 3%.

b. Similarly plant bodies are made up of cellulose, which is composed of carbon, hydrogen and oxygen.

iii. Essential for existence life:

Life owes to non-metals as without O_2 and CO_2 (essential gases for respiration of animals and plants respectively), life would not have been possible. In fact, these gases are essential for the existence of life.

iv. Maintenance of life:

All eatables like carbohydrates, proteins, fats, vitamins, water, milk etc which are necessary for the growth and development of body that are made up of non-metals; carbon, hydrogen and oxygen. Its shows non- metals playa vital role for the maintenance of life

v. Survival of life:

The essential compound for the survival of life of both animals and plants is water, which is made up of non-metals. Water is not only the major part by mass of animals and plants bodies, but it is also essential to maintain the life. We can survive without water for days but not for a long period; its shortage may cause death.

vi. Safety of life:

Non-metal nitrogen, which is 78% in atmosphere, is necessary for the safety of life on earth. It controls the fire and combustion processes, otherwise all the things around us could burn with a single flame.

vii. Communication in life:

Non-metals are playing essential role for the communication in life. All fossil fuels which are major source of energy: coal, petroleum and gas are made up of carbon and hydrogen. Even the essential component of combustion of fossil fuels, oxygen is also a non-metal.

viii. Clothing:

Non-metals protect us in a way. The clothes we wear are made of cellulose (natural fiber) or polymer (synthetic fiber).

ix. Manufacture of industrial goods:

In addition to all these, other items used in daily life such as wooden or plastic furniture, plastic sheets and bags, plastic pipes and utensils are made of non-metallic elements.
Even all the pesticides, insecticides, fungicides and germicides consist of non-metals a major constituent.

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com

Unit 8: Chemical Reactivity
Extra MCQ's

1. **Non-metals are essential**
(a) for the maintenance (b) for the existence
(c) for the safety of life (d) all of these
2. **It has the highest percentage in earth's crust and oceans.**
(a) calcium (b) carbon (c) oxygen (d) nitrogen
3. **The electronegative and non-metallic character from top to bottom**
(a) increases (b) decreases (c) remain same (d) stable
4. **Chlorine only reacts with methane in**
(a) darkness (b) sunlight (c) yellow light (d) screened light
5. **Chlorine has colour:**
(a) pale yellow (b) greenish yellow (c) reddish brown (d) purple black
6. **These are highly oxidizing agents**
(a) alkali metals (b) alkaline earth metals
(c) transition metals (d) halogens
7. **The highest electronegative element in periodic table is**
(a) fluorine (b) chlorine (c) bromine (d) iodine
8. **The electronic configuration of halogens is**
(a) $ns^2\ np^5$ (b) $ns^2\ np^3$ (c) $ns^3\ np^5$ (d) $ns^2\ np^2$
9. **Non-metals are electronegative in nature and form oxides.**
(a) acidic (b) basic (c) neutral (d) suboxide
10. **It is used to make jewelry items because of its unique characteristics like colour beauty, strength, flexibility and resistance to tarnish.**
(a) gold (b) platinum (c) silver (d) copper
11. **White gold is an alloy of**
(a) gold + palladium + zinc (b) palladium + silver + nickel
(c) gold + silver + copper (d) palladium + nickel + zinc
12. **The alkaline earth metals are smaller in size and have more**
(a) ionization energy (b) electron affinity (c) nuclear change (d) electropositive
13. **All alkali metals have the largest size and the lowest in their respective periods.**
(a) electro negativity (b) ionization energy (c) electron affinity (d) electro positivity
14. **Metals have the tendency to lose their valence electron. This property of a metal is termed as**
(a) electro negativity (b) electro positivity (c) electron affinity (d) ionization power
15. **Copper, mercury, silver and gold are the examples of metals**
(a) very reactive (b) moderately (c) least reactive (d) none of these
16. **Cation in formed, when an element – electron to its outermost shell**
(a) loses (b) gains (c) donates (d) shares
17. **Metals which are easily oxidized are said**
(a) negative metals (b) state metals (c) reactive metals (d) non-reactive metals
18. **A metal in a compound always exists in which oxidation sate**
(a) negative (b) positive (c) neutral (d) zero
19. **Ionization energy of sodium is less than**

(a) aluminum (b) magnesium (c) copper (d) all of these

20. All metals are solids except
(a) sodium (b) magnesium (c) mercury (d) gold

21. The most valuable metal among the following is
(a) Gold (b) Uranium
(c) Osmium (d) Rubidium

22. The heaviest metal is
(a) Uranium (b) Gold
(c) Osmium (d) Calcium

23. Lithium has density
(a) $0.53gcm^{-3}$ (b) $1.53gcm^{-3}$
(c) $15.3gcm^{-3}$ (d) $3.5gcm^{3}$

24. Mark which one is non-metal?
(a) Sodium (b) Calcium
(c) Nitrogen (d) Mercury

25. Which one of the following will not react with HCl?
(a) Carbon (b) Silver
(c) Zinc (d) Copper

26. Sodium does not react with
(a) Carbon (b) Nitrogen
(c) Hydrogen (d) Both a and b

27. Which metal burns with golden yellow flame
(a) Calcium (b) Barium
(c) Sodium (d) Potassium

28. Which metal is used in Thermite process to ignite aluminum powder
(a) Na (b) Mg
(c) Ca (d) Be

29. Silver is get tarnished by
(a) Atmosphere oxygen (b) Nitrogen
(c) Hydrogen sulphide (d) Carbon dioxide

30. Which metal is used for making mirrors?
(a) Lead (b) Iron
(c) Silver (d) Lithium

31. Which one of the following metal has yellow colour
(a) Lead (b) Gold
(c) Iron (d) Potassium

32. Gold is always alloyed with one, among the following metals
(a) Sodium (b) Mercury
(c) Copper (d) Calcium

33. Sodium is used for the refining of which metal
(a) U (b) Zn
(c) Ti (d) Zr

34. How many series of d-block elements are there in the periodic
(a) Three (b) Four
(c) Five (d) Two

35. Three transition elements Cu, Ag and Au constitute group number
(a) 9 (b) 10
(c) 11 (d) 12

36. The least electronegative element among the halogens is
(a) Fluorine (b) Chlorine (c) Bromine (d) Iodine

37. Which of the following halogen has pale yellow colour?
(a) F_2 (b) Cl_2 (c) Br_2 (d) I_2

38. Which one of the following makes covalent bond with halogens
(a) Na (b) K (c) O (d) Mg

39. Identify among the following which one is semimetal
(a) Lead (b) Zinc (c) Silicon (d) Galium

40. The general electronic valence shell configuration of alkali metals is
(a) ns^2 (b) ns^1 (c) ns^2, np^1 (d) ns^2, np^2

41. Hydrogen is released when water reacts with
(a) Na (b) Mg (c) K (d) All of them

42. Which metal releases electron from its outermost shell the most easily
(a) Na (b) K (c) Mg (d) Ca

43. Which one among the halogens has least affinity with hydrogen?
(a) F_2 (b) Cl_2 (c) Br_2 (d) I_2

44. In diffused sun lights chlorine reacts with methane to form
(a) CH_3Cl (b) CH_2Cl_2 (c) CCL_4 (d) All of above

45. In the earth crust the highest %age of oxygen is
(a) 47% (b) 86% (c) 90% (d) 24%

46. The %age of oxygen in oceans is:
(a) 47% (b) 86% (c) 90% (d) 24%

47. The %age of oxygens in air is:
(a) 21% (b) 24% (c) 26% (d) 30%

48. Which one of the following will form amphoteric oxides?
(a) Na (b) N (c) Si (d) S

49. Which halogen reacts with water in dark and cold state
(a) F_2 (b) Cl_2 (c) Br_2 (d) I_2

50. Indicate the least reactive metal among the following
(a) Gold (b) Zinc (c) Iron (d) Tin

ANSWER KEY

1	d	11	d	21	b	31	b	41	d
2	c	12	b	22	c	32	c	42	b
3	b	13	a	23	a	33	c	43	d
4	b	14	b	24	c	34	a	44	d
5	b	15	c	25	a	35	c	45	a
6	d	16	a	26	d	36	d	46	b
7	a	17	c	27	c	37	a	47	a
8	c	18	b	28	b	38	c	48	c
9	a	19	d	29	c	39	d	49	a
10	b	20	c	30	c	40	b	50	a

Last Updated: November 2020

Report any mistake at freeilm786@gmail.com